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ABSTRACT

ChemSource is designed as a strategy to help preservice and inservice high school chemistry teachers promote student learning more effectively. Its major premise is that well-designed laboratory investigations are an important avenue for cultivating student interest, engagement, and meaningful learning in chemistry. The SourceBook component of ChemSource is a resource containing 36 content modules and additional supporting materials which aim at providing specific teaching tips and the best available instructional ideas and information gleaned from experienced, successful chemistry teachers. Each content module features: Content-in-a-Nutshell; Laboratory Activities; Demonstrations; Common Misconceptions; Humor; Applications and Implications; Analogies; Reference Resources; History; and Media. General materials on pedagogy, classroom management, safety, teaching chemistry to the disabled, and other useful resources and tabulations precede and follow the content modules. Volume 3 includes the following: Nuclear Chemistry; Organic Chemistry; Oxidation-Reduction; Periodicity; Photochemistry; Polymers; Solubility and Precipitation; Inorganic Qualitative Analysis; Rates of Reaction; Chemistry of Rocks, Minerals, and Gems; and Simple Chemical Reactions. (JRH)

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SourceBook

Version 2.0 1994

ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

VOLUME 3

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SE 055-642

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- Preface and Acknowledgements (PREF)
- Chemical Pedagogy (PEDA)
- User's Guide (USER)
- Introductory: General Resource (INTR)
- Acids and Bases (ACID)
- Alkali Metals (ALKA)
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- Biogeochemical Cycles (BICY)
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A SourceBook Module

Version 2.0 1994

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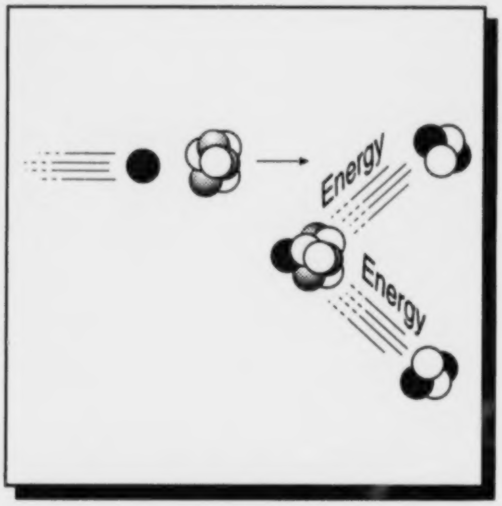
NUCLEAR CHEMISTRY



ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

When the universe came into being, the elements formed by a coming together of protons and neutrons, accompanied by an enormous emission of energy. Our sun and other stars continue to fuel nuclear reactions and emit their associated energy.

The nuclei of atoms contain protons and neutrons held together by strong binding forces. Over one hundred kinds of subatomic particles of nuclear origin have been identified. The so-called fundamental particles, protons and neutrons, are made up of more basic particles called quarks.

Chemical reactions take place as a result of changes in the higher electron energy levels in the atom. If these changes are spontaneous they lead to greater stability of products with respect to reactants. Nuclear reactions are the result of changes within the nucleus of unstable atoms. An unstable nucleus, with a very high or very low neutron to proton ratio, may lead to natural radioactivity with alpha, beta, or gamma emission. This emission cannot be delayed or stopped by any external activity nor does it depend upon whether the element is free or in a compound. Radioactivity results in transmutation of an unstable nucleus to a stable nucleus. Stable nuclei with a favorable neutron to proton ratio do not spontaneously decay. (See Decay Pattern of $^{235}_{92}U$ in *Appendix*.)

The splitting of very heavy nuclei is called nuclear fission. When light nuclei combine, the reaction is referred to as nuclear fusion. These changes lead to more stable nuclei with the release of great amounts of energy. We make use of natural and induced nuclear reactions in power plants, nuclear weapons, medical treatment, and consumer products.

CENTRAL CONCEPTS

1. The nucleus is composed of protons (with a +1 charge) and neutrons (no charge), which are in turn composed of smaller particles called quarks.
2. There is a critical relationship between the number of protons and neutrons in a nucleus, determining its stability.
3. An unstable nucleus commonly decays by one of several radioactive modes: alpha (α), beta (β), and/or gamma (γ) emission.
4. The positively charged particles are held together in the nucleus due to very strong binding energy.
5. The energy of nuclear reactions is significantly larger than that associated with physical change or chemical reactions.
6. The relative instability of a nucleus is indicated by a characteristic rate of decay, measured by its half-life.
7. There are many applications of radioisotopes in medicine, commercial processes and research.

RELATED CONCEPTS

1. An understanding of the atomic model (see *Atomic Structure* module).
2. A qualitative understanding that energy changes are involved in all reactions.
3. Thermochemical calculations.

RELATED SKILLS

1. Ability to write and balance chemical equations.
2. Ability to perform stoichiometric calculations.

PERFORMANCE OBJECTIVES

After completing their study of nuclear chemistry, students should be able to:

1. balance a nuclear equation.
2. compare the relative amounts of energy released from physical, chemical, and nuclear reactions.
3. calculate the quantity of material remaining after a specified number of half-lives.
4. identify the relative benefits and hazards of nuclear reactions.

Concept/Skills Development

LABORATORY ACTIVITY AND DEMONSTRATIONS

This topic does not have any feasible, applicable laboratory activities associated with it. If, by some chance, your high school owns a Geiger counter (try the physics teacher) or cloud chamber and a source of α or β emission, these items can be used for demonstrations. A sealed β source can be used to expose Polaroid™ film. Half life can be demonstrated by starting out with a known number of candy corn or other object and diminishing by half each time. Dismissing a class by coin toss can simulate the decay process (probability).

GROUP AND DISCUSSION ACTIVITIES

1. Relative energy calculations comparing the output of physical, chemical, and nuclear processes (see *Appendix*). Discuss the abundance of raw materials and compare that with the efficiency of energy conversion.
2. Have students understand the effects of a chain reaction. These effects can be illustrated by:
   a. generating a "picture in the mind" of a room full of mousetraps, each one with a ping-pong ball poised on the bait end of the trap. Now imagine tossing one ping-pong ball into the room and triggering the chain reaction.
   b. having students calculate how many cycles of letters it would take, in a high school of 1000 students, to make sure everyone received a chain letter.
3. Two excellent activities (half-life simulation and understanding isotopes) are included in the *ChemCom* unit on radioactive decay. (See *References* for publisher information.)

TIPS FOR THE TEACHER

Language of Chemistry

1. **Glossary**

alpha particles(a,) $^{4}_{2}He$ a nuclear particle containing two protons and two neutrons with a mass number of four (4.0026 u). An alpha particle is identical to a helium nucleus. Alpha particles are most often emitted from very heavy nuclei (*i.e.*, $Z \geq 83$).

beta particle (b, $^{0}_{-1}e$) a nuclear particle of zero mass number (0.0005486 u) and unit negative charge. A beta particle is an electron of nuclear origin. Beta particles may be emitted from both heavy and light radioactive nuclei.

chain reaction a reaction in which one of the products can also serve as one of the reactants. Such a reaction is self-sustaining, once begun. In nuclear fission, the chain reaction is started with neutrons that split uranium nuclei, which, in turn release more neutrons to sustain the reaction.

control rods cylinders usually made of neutron-absorbing material, such as boron or cadmium, which may be lowered into a nuclear fission reactor to control the speed of the reaction by slowing, or halting the chain process.

critical mass the minimum mass of radioactive material needed to sustain fission.

fission the process whereby certain heavy nuclei (*e.g.*, $^{235}_{92}U$) break apart into two approximately equal smaller, more stable, nuclei called "daughters." Fission is usually initiated by bombarding heavy nuclei with slow moving neutrons. In the process, additional neutrons are released, which can cause the fission of yet more nuclei, thus sustaining a chain reaction. Fission processes release large quantities of energy.

fusion the process in which two small nuclei (usually H or He) combine to form a larger, more stable, nucleus. The heat and light of the sun are generated by the fusion of hydrogen isotopes to form helium nuclei.

gamma rays (γ) high energy electromagnetic radiation. The wavelength of gamma rays is approximately 10^{-12} m (refer to a standard chemistry textbook for electromagnetic spectrum diagram). Both alpha and beta decay are usually accompanied by gamma emission.

half-life ($t_{1/2}$) time required for one-half of a given amount of a radioactive substance to decay.

isotopes nuclei of an element having the same number of protons but different number of neutrons. For any given element, some isotopes may be unstable (and therefore undergo nuclear decay), while others may be stable.

moderator a substance that slows down neutrons in a fission reactor.

nuclear decay the general term for any of several processes (alpha emission, beta emission, fission, *etc.*) that an unstable nucleus can undergo in order to achieve stability.

nuclear reactor vessel where controlled nuclear fission takes place.

positron (e^+, $^{0}_{+1}e$) positively charged particle with the same mass as an electron.

radioactivity the generic term for alpha, beta or gamma emission.

transuranium elements the elements heavier than uranium (U), therefore those elements with Z > 92.

units of radiation various types of units exist to characterize ionizing radiation (alpha, beta, and gamma). Each is a slightly different way to measure the radiation.

a. A **rad** , or **r**adiation **a**bsorbed **d**ose, corresponds to 10^{-5} J of energy absorbed per gram of body weight. Not all forms of ionizing radiation have the same biological effect, so the **RBE**, **r**adiation **b**iological **e**ffectiveness, is a correction factor. For instance, X-rays and gamma rays are assigned an RBE of 1, while alpha particles have an RBE of 10, which means alpha particles can cause more ionization and therefore more damage to the body. This term can be misleading because alpha particles external to the body are stopped by our skin, or even a piece of paper; they have little penetrating power. Gamma rays, on the other hand can penetrate soft body parts.

b. The **rem** , **r**oentgen **e**quivalent **m**an, is calculated by multiplying the absorbed dose in rads by the radiation biological effectiveness.

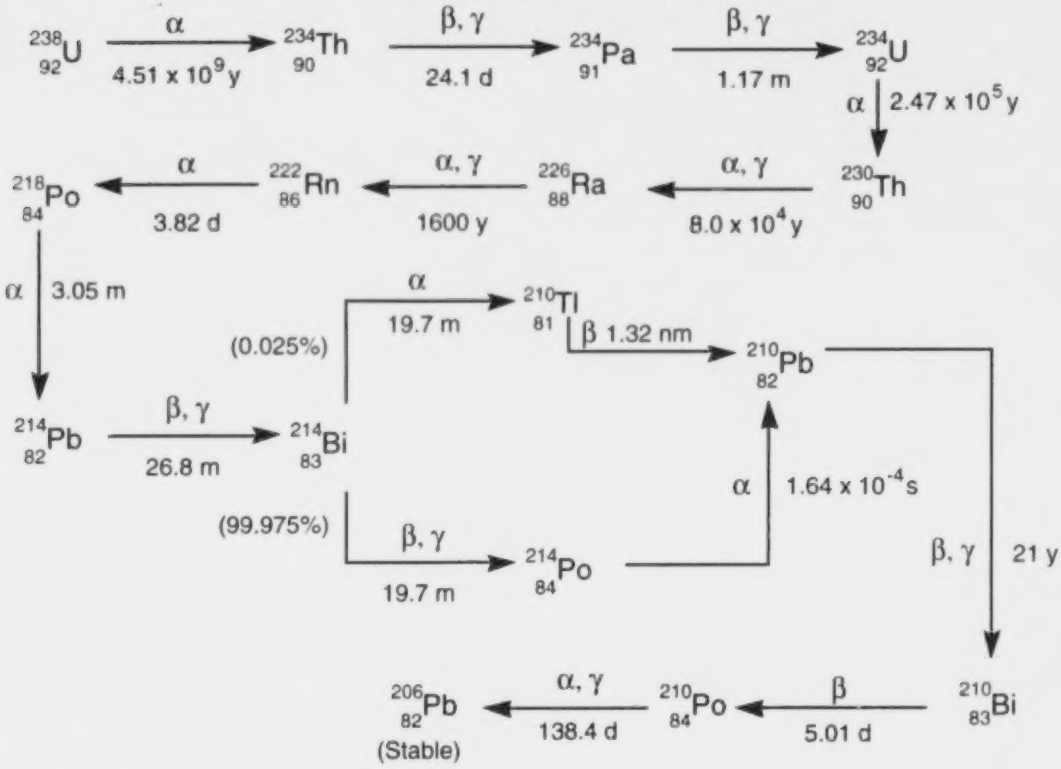
$$\text{rem} = \text{rad} \times \text{RBE}$$

c. The **activity** of a radioactive sample is usually measured in disintegrations per second, or in curies, where

$$1 \text{ curie (Ci)} = 3.7 \times 10^{10} \text{ disintegrations/s}$$

d. A **roentgen** is a unit of radiation dose that produces one electrostatic unit of electricity per cubic centimeter of dry air at 0 °C and 1 atm pressure.

2. Decay Pattern (of Uranium-238) and Half-Life:



$^{238}_{92}U$
α
4.51 x 10⁹ y
$^{234}_{90}Th$
β, γ
24.1 d
$^{234}_{91}Pa$
β, γ
1.17 m
$^{234}_{92}U$
α
2.47 x 10⁵ y
$^{230}_{90}Th$
α, γ
8.0 x 10⁴ y
$^{226}_{88}Ra$
α, γ
1600 y
$^{222}_{86}Rn$
α
3.82 d
$^{218}_{84}Po$
α
3.05 m
$^{214}_{82}Pb$
β, γ
26.8 m
$^{214}_{83}Bi$
(0.025%)
α
19.7 m
$^{210}_{81}Tl$
β 1.32 nm
$^{210}_{82}Pb$
(99.975%)
β, γ
19.7 m
$^{214}_{84}Po$
α
1.64 x 10⁻⁴ s
β, γ
21 y
$^{210}_{83}Bi$
β
5.01 d
$^{210}_{84}Po$
α, γ
138.4 d
$^{206}_{82}Pb$
(Stable)

y = years, d = days, m = minutes, s = seconds

Common Student Misconceptions

1. **"Irradiated food is radioactive."**

   Although some foods have an extended shelf-life after radiation with gamma rays, irradiation only kills the microorganisms in the food. This process does not make the food radioactive.

2. **"All radiation is harmful."**

   The physical harm caused by radiation varies greatly, depending on the quantity received, what type it is (gamma rays can be more harmful, in general, than alpha particles), whether the radiation is internal or external to the body, *etc.* (see the tables in the *Appendix* for additional comparisons).

3. **"All radiation is man-made."**

   Most radiation to which humans are exposed is natural radiation from cosmic rays or from naturally occurring radioisotopes.

4. **"I can't be hurt by something I can't see or feel."**

   Too much radiation can indeed be harmful. Just because it isn't visible doesn't mean it isn't real.

5. **"All uranium can be used as nuclear fuel."**

   Only the ^{235}U isotope of uranium is fissionable; most uranium ore (>90%) is ^{238}U, which is not fissionable.

6. **"The number of counts from a Geiger counter is equal to the number of nuclei that decay."**

   The Geiger counter detects and records only a small fraction of particles emitted from decaying nuclei. Some reasons for this inability to count all disintegrations are:

   a. Window thickness of the Geiger tube. Not all particles will penetrate the tube.

   b. Geiger tube dead time. While the tube is busy reacting to a particle, other particles go undetected.

   c. Counter geometry. Radioisotopes emit radiation in all directions (4-π geometry). A typical Geiger tube can only monitor a small fraction or small angle of all the radiation direction angles.

   A typical counter only counts the relative number of disintegrations.

Pattern Recognition/Problem Solving

1. To balance nuclear equations, conservation of both mass number and charge on the nucleus must be achieved. **Mass number** (A) is the sum of the numbers of protons and neutrons in the nucleus. It is also the approximate mass of a particle or isotope to the nearest whole atomic mass unit, u. **Charge** is the positive or negative whole units of electrical charge. One unit equals 1.60×10^{-19} C (coulombs).

2. Nuclear Symbols. — $^{A}_{Z}X$

   A = mass number, and Z = positive charge on nucleus or
   A = sum of the protons and neutrons, and Z = number of protons

   e.g., for $^{238}_{92}U$ 238 = mass number, and 92 = charge on nucleus or
   238 = protons + neutrons, and 92 = protons

other:

proton: $^{1}_{1}H$ electron: $^{0}_{-1}e$ neutron: $^{1}_{0}n$

alpha (α) particle: $^{4}_{2}He$ beta (β or β^-) particle: $^{0}_{-1}e$ positron (β^+): $^{0}_{+1}e$

3. Balancing Nuclear Equations. — In a nuclear equation charge and mass must be conserved.

   Examples:

   a. **Alpha radiation** ($^{4}_{2}He$ emitted from nucleus)

$$^{238}_{92}U \rightarrow {}^{4}_{2}He + {}^{A}_{Z}X$$

238 = (4 + A) A must be 234 (conservation of mass number)
92 = (2 + Z) Z must be 90 (conservation of charge)

The new isotope produced is $^{234}_{90}\mathrm{Th}$ (by finding the symbol for Z = 90 in the Periodic Table).

b. **Beta radiation** ($^{0}_{-1}\mathrm{e}$ emitted from nucleus):

$$^{210}_{82}\mathrm{Pb} \rightarrow {}^{0}_{-1}\mathrm{e} + {}^{A}_{Z}\mathrm{X}$$

$210 = (0 + A)$	A must be 210
$82 = (-1 + Z)$	Z must be 83

The new isotope produced is $^{210}_{83}\mathrm{Bi}$.

c. **Fission** (see *Language of Chemistry*):

$$^{235}_{92}\mathrm{U} + {}^{1}_{0}\mathrm{n} \rightarrow {}^{139}_{56}\mathrm{Ba} + {}^{94}_{36}\mathrm{Kr} + 3\,{}^{1}_{0}\mathrm{n} + \text{Energy}$$

$235 + 1 = 139 + 94 + 3(1)$ (conservation of mass number)
$92 + 0 = 56 + 36 + 3(0)$ (conservation of charge)

NOTE: Many other products of uranium fission in addition to barium and krypton are formed.

d. **Fusion** (see *Language of Chemistry*):

$$4\,{}^{1}_{1}\mathrm{H} \rightarrow {}^{4}_{2}\mathrm{He} + 2\,{}^{0}_{1}\mathrm{e} + \text{Energy}$$

$4(1) = 4 + 2(0)$ (conservation of mass number)
$4(1) = 2 + 2(1)$ (conservation of charge)

4. Student Worksheet: Writing Nuclear Equations

Part A. Complete each of the following by identifying $^{A}_{Z}\mathrm{X}$.

1. $^{235}_{92}\mathrm{U} + {}^{1}_{0}\mathrm{n} \rightarrow {}^{A}_{Z}\mathrm{X}$
2. $^{222}_{88}\mathrm{Ra} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{A}_{Z}\mathrm{X}$
3. $^{A}_{Z}\mathrm{X} + {}^{0}_{-1}\mathrm{e} \rightarrow {}^{54}_{24}\mathrm{Cr}$
4. $^{54}_{26}\mathrm{Fe} + {}^{1}_{0}\mathrm{n} \rightarrow {}^{1}_{1}\mathrm{H} + {}^{A}_{Z}\mathrm{X}$
5. $^{238}_{92}\mathrm{U} + {}^{12}_{6}\mathrm{C} \rightarrow {}^{246}_{98}\mathrm{Cf} + 4\,{}^{A}_{Z}\mathrm{X}$
6. $^{A}_{Z}\mathrm{X} + {}^{4}_{2}\mathrm{He} \rightarrow {}^{245}_{97}\mathrm{Bk} + {}^{1}_{1}\mathrm{H} + 2\,{}^{1}_{0}\mathrm{n}$
7. $^{235}_{92}\mathrm{U} + {}^{1}_{0}\mathrm{n} \rightarrow {}^{139}_{56}\mathrm{Ba} + {}^{A}_{Z}\mathrm{X} + 3\,{}^{1}_{0}\mathrm{n}$
8. $^{2}_{1}\mathrm{H} + {}^{3}_{1}\mathrm{H} \rightarrow {}^{A}_{Z}\mathrm{X} + {}^{1}_{0}\mathrm{n}$

Part B. Write balanced equations for each of these reactions:

1. Alpha emission of ${}^{214}_{84}\text{Po}$
2. Beta emission of ${}^{210}_{83}\text{Bi}$
3. Neutron emission of ${}^{107}_{47}\text{Ag}$
4. Positron emission of ${}^{70}_{33}\text{As}$
5. Neutron emission of ${}^{88}_{35}\text{Br}$
6. Electron absorption by ${}^{116}_{51}\text{Sb}$
7. Proton emission of ${}^{41}_{19}\text{K}$
8. Alpha absorption by ${}^{14}_{7}\text{N}$ with neutron emission

Worksheet Answers

Part A

1. ${}^{236}_{92}\text{U}$
2. ${}^{218}_{86}\text{Rn}$
3. ${}^{54}_{25}\text{Mn}$
4. ${}^{54}_{25}\text{Mn}$
5. ${}^{1}_{0}\text{n}$
6. ${}^{244}_{96}\text{Cm}$
7. ${}^{94}_{36}\text{Kr}$
8. ${}^{4}_{2}\text{He}$

Part B

1. ${}^{214}_{84}\text{Po} \rightarrow {}^{4}_{2}\text{He} + {}^{210}_{82}\text{Pb}$
2. ${}^{210}_{83}\text{Bi} \rightarrow {}^{0}_{-1}\text{e} + {}^{210}_{84}\text{Po}$
3. ${}^{107}_{47}\text{Ag} \rightarrow {}^{106}_{47}\text{Ag} + {}^{1}_{0}\text{n}$
4. ${}^{70}_{33}\text{As} \rightarrow {}^{0}_{+1}\text{e} + {}^{70}_{32}\text{Ge}$
5. ${}^{88}_{35}\text{Br} \rightarrow {}^{1}_{0}\text{n} + {}^{87}_{35}\text{Br}$
6. ${}^{116}_{51}\text{Sb} + {}^{0}_{-1}\text{e} \rightarrow {}^{116}_{50}\text{Sn}$
7. ${}^{41}_{19}\text{K} \rightarrow {}^{1}_{1}\text{H} + {}^{40}_{18}\text{Ar}$
8. ${}^{14}_{7}\text{N} + {}^{4}_{2}\text{He} \rightarrow {}^{17}_{9}\text{F} + {}^{1}_{0}\text{n}$

5. Half-Life of Radon (see *Appendix*).

HISTORY: ON THE HUMAN SIDE

1834	Michael Faraday—electrolysis experiments suggested electrical nature of matter.
1895	Wilhelm Roentgen—discovered X-rays when cathode rays strike the anode.
1896	Henri Becquerel—discovered "uranic rays" and radioactivity.
1896	Marie (Marya Sklodowska) and Pierre Curie—discovered that radiation is a property of the atom, and not due to chemical reaction. (Marie named this property *radioactivity.*)
1897	Joseph J. Thomson—discovered the electron through Crookes tube experiments.
1898	Marie and Pierre Curie—discovered the radioactive elements polonium and radium.
1899	Ernest Rutherford—discovered alpha and beta particles.
1900	Paul Villard—discovered gamma rays.
1903	Ernest Rutherford and Frederick Soddy—established the laws of radioactive decay and transformation.
1910	Frederick Soddy—proposed the isotope concept to explain the existence of more than one atomic weight of radioelements.
1911	Ernest Rutherford—used alpha particles to explore gold foil; discovered the nucleus and the proton; proposed the nuclear theory of the atom.
1919	Ernest Rutherford—announced the first artificial transmutation of atoms.
1932	James Chadwick—discovered the neutron by alpha particle bombardment of Beryllium.
1934	Frederic Joliot and Irene Joliot Curie—produced the first artificial radioisotope.
1938	Otto Hahn, Fritz Strassmann, Lise Meitner, and Otto Frisch—discovered nuclear fission of uranium-235 by neutron bombardment.
1940	Edwin M. McMillan and Philip H. Abelson—discovered the first transuranium element, neptunium, by neutron irradiation of uranium in a cyclotron.
1941	Glenn T. Seaborg, Edwin M. McMillan, Joseph W. Kennedy and Arthur C. Wahl—announced discovery of plutonium from beta particle emission of neptunium.
1942	Enrico Fermi—produced the first nuclear fission chain-reaction.
1944	Glenn T. Seaborg—proposed a new format for the periodic table to show that a new actinide series of 14 elements would fall below and be analogous to the 14 lanthanide-series elements.
1964	Murray Gell-Mann hypothesized that quarks are the fundamental particles that make up all known subatomic particles except leptons.

HUMOR: ON THE FUN SIDE

1. **ELEMENTARY GOSPEL** *(CHEM 13 NEWS, November 1975, p. 977)*

   And the Lord said unto uranium, "You shall lead a degenerate life," and it was lead.

2. Word search (see *Appendix* for master copy)

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F J V F C M O D E R A T O R R M N
C U B F Q W Z Q E L V Z A A N A S
D Y S E Y Q S B S F Z D P C I Q S
H B I I H P K D K U I P Y K S Z A
B O R F O C R J P O V L V X Z H M
W C C Y Q N I Q A N A B F X P E L
T A T E B L P C H O V U U L D W A
F Z Y M K G T P G I J V A U A G C
D C M S J I A K C S G A V B V H I
B I S D V G H G I S O T O P E S T
M D C I W T I E O I W V M V T V I
X C T X D A R O A F N L B N T E R
Z Y Z D T O B W O E T A V L U T C
```

   Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

   1. Nuclear particle containing two neutrons and two protons.
   2. Particle that is an electron of nuclear origin.
   3. Process of heavy nuclei breaking down into smaller nuclei called "daughters."
   4. Process wherein two small nuclei combine to form a larger, more stable, nucleus.
   5. Time required for one-half of a given amount of a radioactive substance to decay. (2 words)
   6. Nuclei of an element having the same number of protons but different numbers of neutrons.
   7. Generic term for the emission of alpha, beta, or gamma rays.
   8. Radiation absorbed dose, 10^{-5} joules/gram of body mass. (Abbrev.)
   9. Describing the minimum mass of radioactive material needed to sustain fission.
   10. Substance that slows down neutrons in a fission reactor.

   Answers: 1. ALPHA 2. BETA 3. FISSION 4. FUSION 5. HALF-LIFE 6. ISOTOPES 7. RADIOACTIVITY 8. RAD 9. CRITICAL 10. MODERATOR

3. See cartoons at end of module.

MEDIA

1. CHEMStudy Film *Transuranium Elements*, available from Ward's Natural Science Establishment, Inc., P.O. Box 92912, Rochester, NY 14692-9978, (716) 334-6174.

2. *The Discovery of Radioactivity*, International Film Bureau, Inc., 332 South Michigan Avenue, Chicago, IL 60604-4382.
3. *Ring of Truth*, episode on energy is pertinent to this topic. Videotapes are available from PBS Home Video (Phone: 1-800-424-7963).
4. *Nova*, PBS series on the Chernobyl nuclear accident. PBS Video, 1320 Bradford Place, Alexandria, VA 22314, (703) 734-5010.
5. *Radiation and Your Environment*, from EME, P.O. Box 2805, Danbury, CT 06813-2805.
6. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).
    a. For the Apple II computer: AP 1001, AP 1002.
    c. For IBM PCs and PC-compatibles: PC 3801.

EQUIPMENT AND INSTRUMENTATION

If your school already owns a Geiger counter you may elect to show your students the radiation from a smoke detector, a lantern mantle, some watch dials, rock samples, *etc.*

Links/Connections

WITHIN CHEMISTRY

1. The number of protons and neutrons determine variations in physical properties of elements observed in isotopes. Naturally radioactive elements decay spontaneously. X-rays can be produced when metallic anodes are bombarded with electrons in a cathode ray tube. (*Nuclear Chemistry* topics may be included with the study of *Atomic Structure*.) The discovery of fundamental particles may be used to develop the atomic model. Radioactive decay and half-life may be used in kinetics to illustrate first-order rate reactions (see *Rates of Reactions* module).
2. Tritium is a naturally occurring radioactive isotope of hydrogen. It has two neutrons and one proton in its nucleus, exists in the diatomic state, and constitutes a very small percentage of the hydrogen content of water. Since it has a half-life of 12.3 years, it can be used for dating rather "young" samples. For example, after a wine is bottled, it becomes a closed system; it can no longer exchange substances with the outside world, and its tritium content will decay at the rate dictated by its half-life. Meanwhile, the naturally occurring level of tritium worldwide remains relatively constant. The tritium level of the wine can be compared to that of ordinary water to determine the wine's age. For example, if a port wine is tested for tritium content along with ordinary water, and it is found that its tritium level is only 30.6% of that of the tritium activity of water, then one can infer that the wine is 21 years old.

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

Nuclear physicists concentrate on the study and expansion of knowledge of subatomic particles, their structure, and models for the nuclear atom. Therefore, nuclear chemistry is an area with strong overlap with a general physics course. Radioactivity, fission, fusion, and nuclear reactors are also important areas to the nonscientist.

TO THE CONTEMPORARY WORLD

Community

Field trips to: nearby nuclear power plant; nuclear medicine center at most large hospitals.

Societal

1. Americium (Am) is the ionizing source in smoke detectors.
2. Nuclear medicine.
3. Carbon-14 is used to date objects made of wood or other formerly living material. Examples include the Shroud of Turin, Native American artifacts, bones, *etc*.
4. The current clamor over radon gas risks in well-insulated homes built over uranium-rich bedrock.
5. Camping lanterns have a mantle containing some radioactive thorium.
6. At one time, watch dials were painted with a paint containing radium in order to make them glow in the dark. The radium was harmful to the health of the factory workers—in particular the women who hand-painted the dials, frequently licking their small paint brushes. Now, tritium (^{3}H) is used, but due to its much shorter half-life, the face may visibly dim during the life of the watch.
7. Some *EXIT* signs are illuminated by tritium rather than by electricity. Can you think why this is an advantage? *[In a power failure, an electrically powered EXIT sign would not remain lighted.]*

Extensions

Nuclear Chemistry is an excellent section for special topics or term papers. Examples of possible subjects are:

1. Linear accelerators.
2. Radon in homes: kits to measure radon gas can be purchased from Central Scientific Co., 11222 Melrose Avenue, Franklin Park, IL 60131-1364 (1-800-262-3626).
3. Any nuclear medical technique. You might contact a local oncologist.
4. Have students consider the pros and cons of generating electricity by nuclear power as opposed to coal-burning plants, or using wood to heat homes. Consider the bumper sticker *Split Wood, Not Atoms*. You might contact your local power company. Most have a technical public relations office that can provide written information or a guest speaker for your class.

References

Module developed by Joseph Baron, Marvin Blevins, and Barbara Sawrey, the California team.

American Chemical Society. (1993). Nuclear chemistry in our world. *ChemCom*. Kendall/Hunt Publishing, pp. 270-337.

An excellent source of material for this topic.

Cline, B. (1989). *The men who made the new physics.* University of Chicago Press.

This is an easy-to-read historical and personal overview of the chemistry and physics that led to the atomic age. Good level for students.

Gamow, G. (1985). *Thirty years that shook physics.* New York: Dover Publications.

Journal of Chemical Education. (1989). *66,* 362-393.

This issue devoted space to seven articles about the history of nuclear fission, the atomic bomb, and using nuclear chemistry in the curriculum.

Tausta, J. (1989, February). Shrouded in mystery. *ChemMatters,* 8-11.

An article specifically for high school students in which C-14 dating of the Shroud of Turin is discussed.

Zuror, P. (1986, May 5). Food irradiation. *Chemical & Engineering News, 64*(18), 46-56.

Discusses the how and why of food irradiation.

Other written information for educators can be obtained from:

The Nucleus, Inc., 761 Emory Valley Road, P.O. Box 2561, Oak Ridge, TN 37830-2561.

To join the American Nuclear Science Teachers Association, write:

American Nuclear Society, 555 North Kensington Avenue, La Grange Park, IL 60525.

Appendix

- **Transparency Masters**
  1. Student Worksheet: Writing Nuclear Equations
  2. Half-Life of Radon
  3. Common Subatomic Species
  4. Word Search
- **Humor**

Student Worksheet: Writing Nuclear Equations

Part A. Complete each of the following by identifying . ${}^{A}_{Z}X$

1. ${}^{235}_{92}U + {}^{1}_{0}n \rightarrow {}^{A}_{Z}X$
2. ${}^{222}_{88}Ra \rightarrow {}^{4}_{2}He + {}^{A}_{Z}X$
3. ${}^{A}_{Z}X + {}^{0}_{-1}e \rightarrow {}^{54}_{24}Cr$
4. ${}^{54}_{26}Fe + {}^{1}_{0}n \rightarrow {}^{1}_{1}H + {}^{A}_{Z}X$
5. ${}^{238}_{92}U + {}^{12}_{6}C \rightarrow {}^{246}_{98}Cf + 4\,{}^{A}_{Z}X$
6. ${}^{A}_{Z}X + {}^{4}_{2}He \rightarrow {}^{245}_{97}Bk + {}^{1}_{1}H + 2\,{}^{1}_{0}n$
7. ${}^{235}_{92}U + {}^{1}_{0}n \rightarrow {}^{139}_{56}Ba + {}^{A}_{Z}X + 3\,{}^{1}_{0}n$
8. ${}^{2}_{1}H + {}^{3}_{1}H \rightarrow {}^{A}_{Z}X + {}^{1}_{0}n$

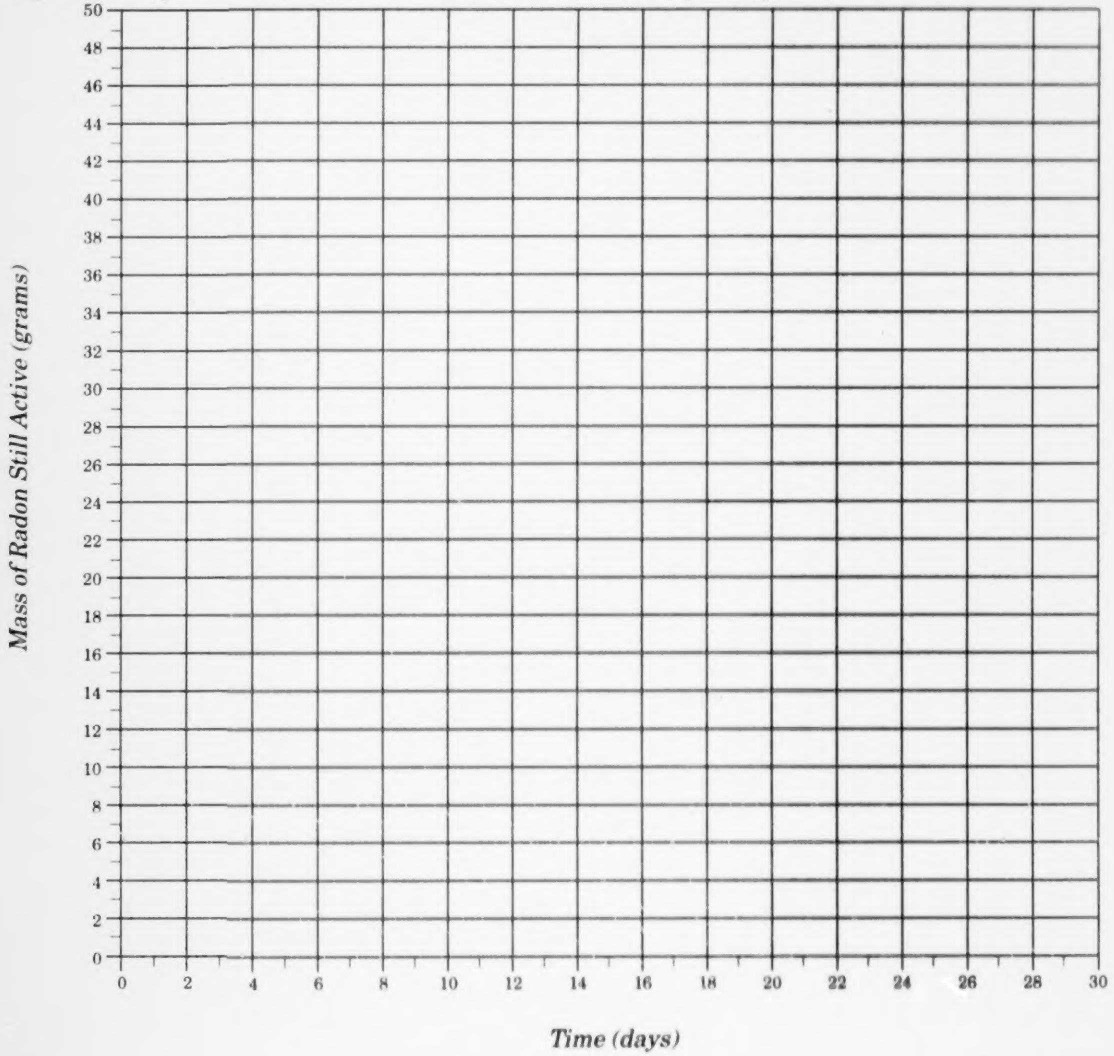
Part B. Write balanced equations for each of these reactions:

1. Alpha emission of ${}^{214}_{84}Po$
2. Beta emission of ${}^{210}_{83}Bi$
3. Neutron emission of ${}^{107}_{47}Ag$
4. Positron emission of ${}^{70}_{33}As$
5. Neutron emission of ${}^{88}_{35}Br$
6. Electron absorption by ${}^{116}_{51}Sb$
7. Proton emission of ${}^{41}_{19}K$
8. Alpha absorption by ${}^{14}_{7}N$ with neutron emission

Half-Life of Radon

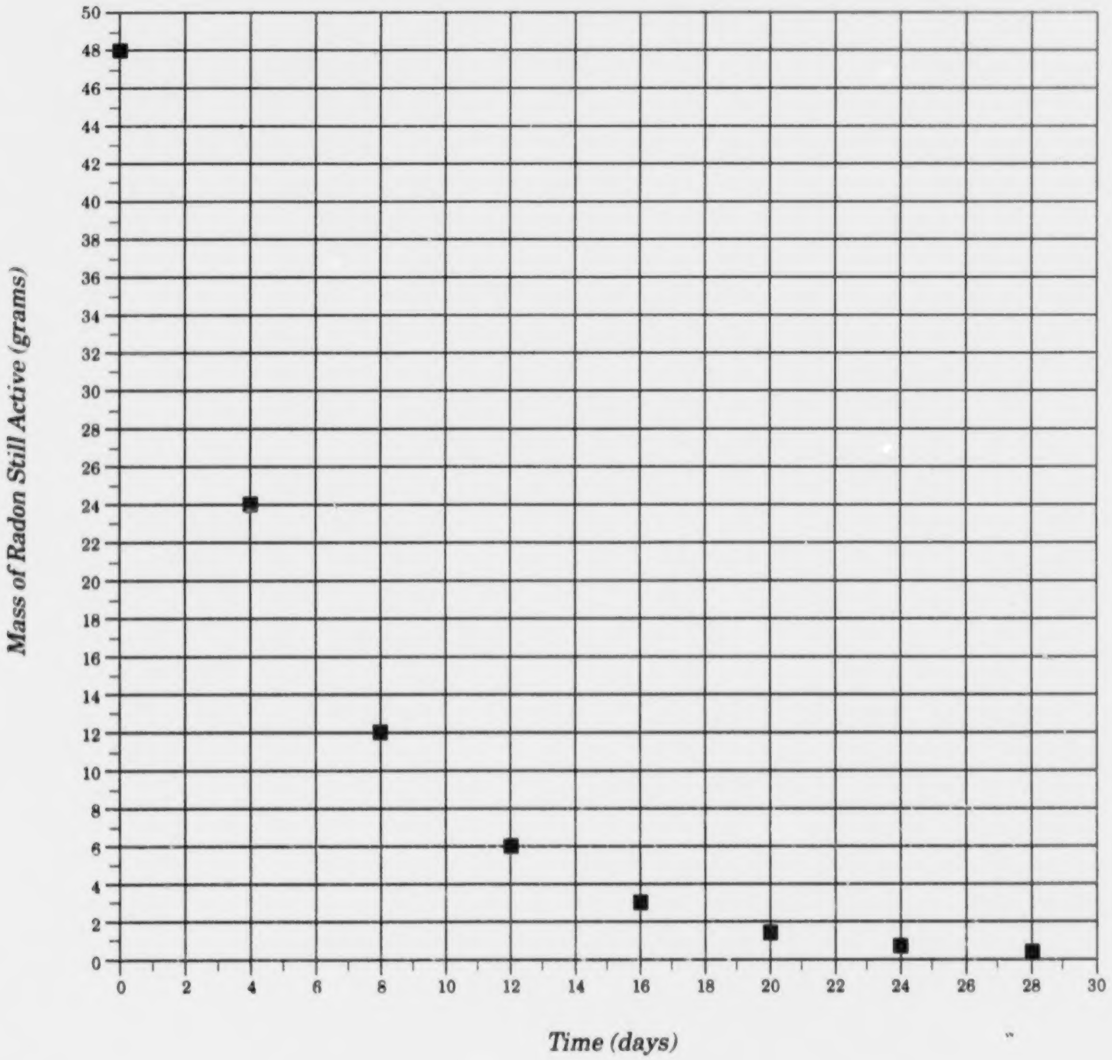
$^{222}_{86}\text{Rn}$ is the first decay product of $^{226}_{88}\text{Ra}$. Radon's activity is much greater than that of its parent. It is the heaviest of all gases (atomic mass = 222), with a density of almost 10 g/L. Like radium, it is hazardous to handle and should be used with adequate protective shielding.

Assume you have an initial sample of radon with a mass of 48 g. Assume the half-life of radon is about four days. On the grid below, plot the amount of radon left after each four-day time period.



Questions

1. How many grams of radon would be present after 4 days?
2. How many grams of radon would be present after 12 days?
3. How many grams of radon would be present after 24 days?
4. If 8 g of radon are left, what is the time elapsed?
5. How many grams of radon would be present after 6 days?



Answers

1. 24 g
2. 6 g
3. 0.75 g
4. 10 days
5. 16 g

Common Subatomic Species

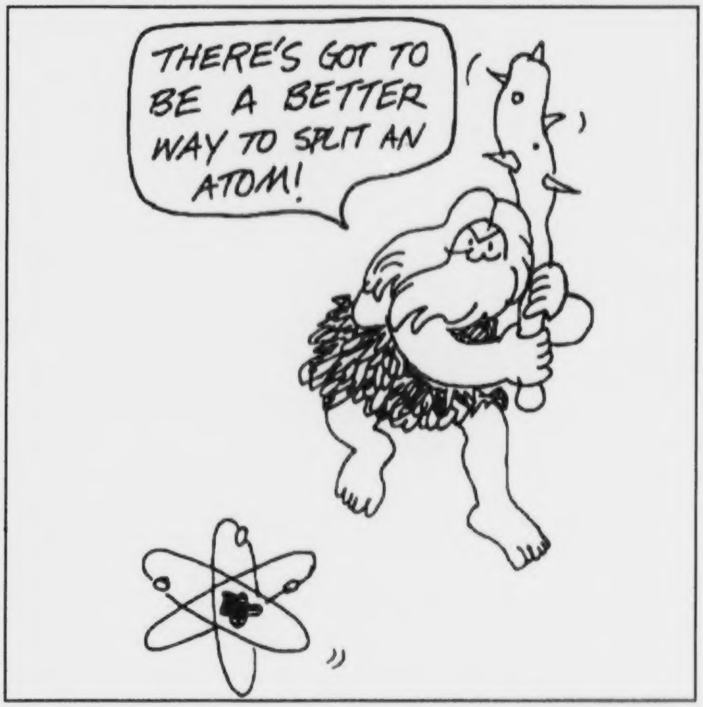
Name	Symbol	Mass(g)	Mass (u)	Mass Number	Charge
electron or beta particle	e^-, β	9.109×10^{-28}	0.0005486	0	–1
proton	p	1.674×10^{-24}	1.0079	1	+1
neutron	n	1.675×10^{-24}	1.0087	1	0
alpha particle or helium nucleus	$^4_2\alpha$, $^4_2\mathrm{He}$	6.647×10^{-24}	4.0026	4	+2
gamma	γ	0	0	0	0
positron	e^+, β^+	9.109×10^{-28}	0.0005486	0	–1

Word Search

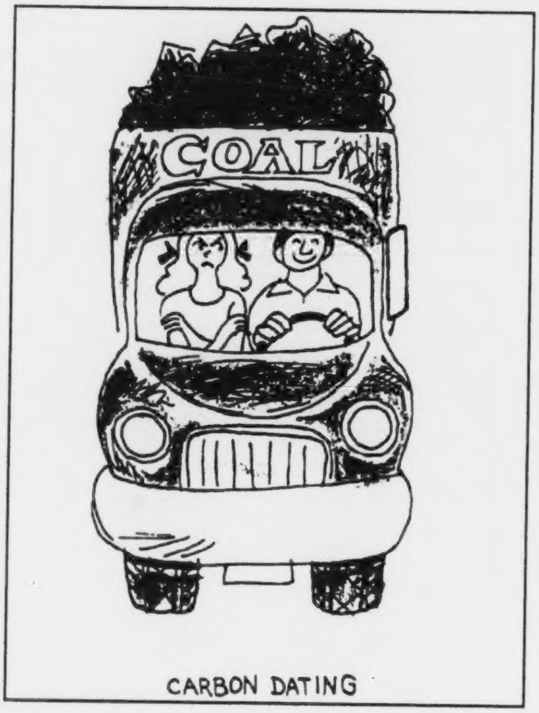
```
F J V F C M O D E R A T O R R M N
C U B F Q W Z Q E L V Z A A N A S
D Y S E Y Q S B S F Z D P C I Q S
H B I I H P K D K U I P Y K S Z A
B O R F O C R J P O V L V X Z H M
W C C Y Q N I Q A N A B F X P E L
T A T E B L P C H O V U U L D W A
F Z Y M K G T P G I J V A U A G C
D C M S J I A K C S G A V B V H I
B I S D V G H G I S O T O P E S T
M D C I W T I E O I W V M V T V I
X C T X D A R O A F N L B N T E R
Z Y Z D T O B W O E T A V L U T C
```

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Nuclear particle containing two neutrons and two protons.
2. Particle that is an electron of nuclear origin.
3. Process of heavy nuclei breaking down into smaller nuclei called "daughters."
4. Process wherein two small nuclei combine to form a larger, more stable, nucleus.
5. Time required for one-half of a given amount of a radioactive substance to decay. (2 words)
6. Nuclei of an element having the same number of protons but different numbers of neutrons.
7. Generic term for the emission of alpha, beta, or gamma rays.
8. Radiation absorbed dose, 10^{-5} joules/gram of body mass. (Abbrev.)
9. Describing the minimum mass of radioactive material needed to sustain fission.
10. Substance that slows down neutrons in a fission reactor.


THERE'S GOT TO BE A BETTER WAY TO SPLIT AN ATOM!

Used by permission of Veritex Publishing Co.


COAL
CARBON DATING

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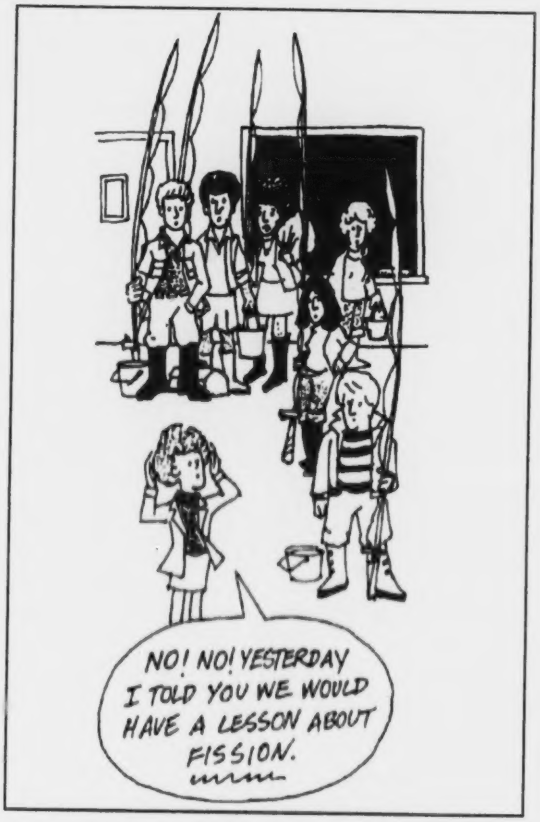
NO! NO! YESTERDAY I TOLD YOU WE WOULD HAVE A LESSON ABOUT FISSION.

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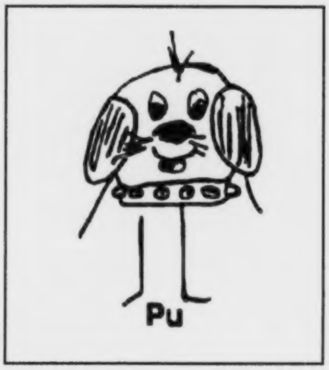

CHEM 13 NEWS, May & September 1986, p. 5, 13.
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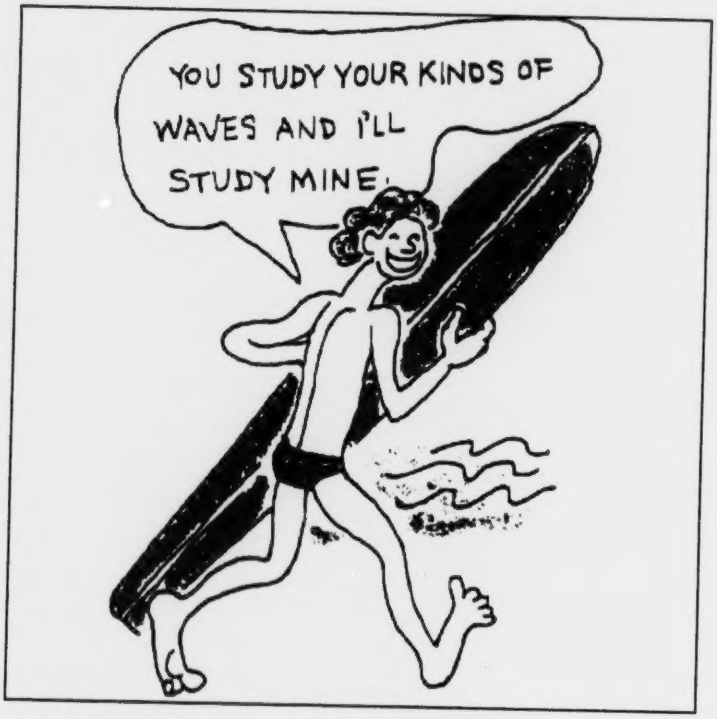
YOU STUDY YOUR KINDS OF WAVES AND I'LL STUDY MINE.

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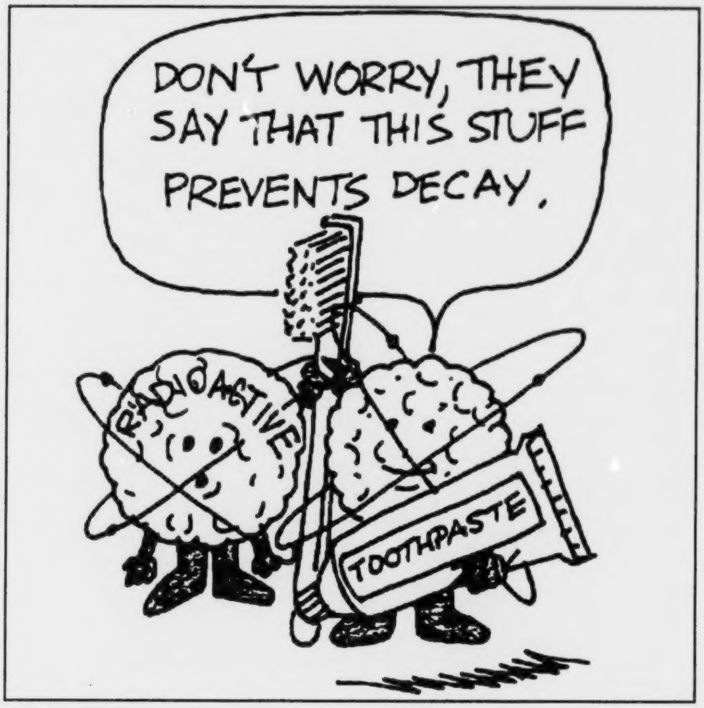

DON'T WORRY, THEY SAY THAT THIS STUFF PREVENTS DECAY.
RADIOACTIVE
TOOTHPASTE

Used by permission of Veritex Publishing Co.

I THINK WE'VE HAD ENOUGH! LET'S SPLIT!
n
n
n
n

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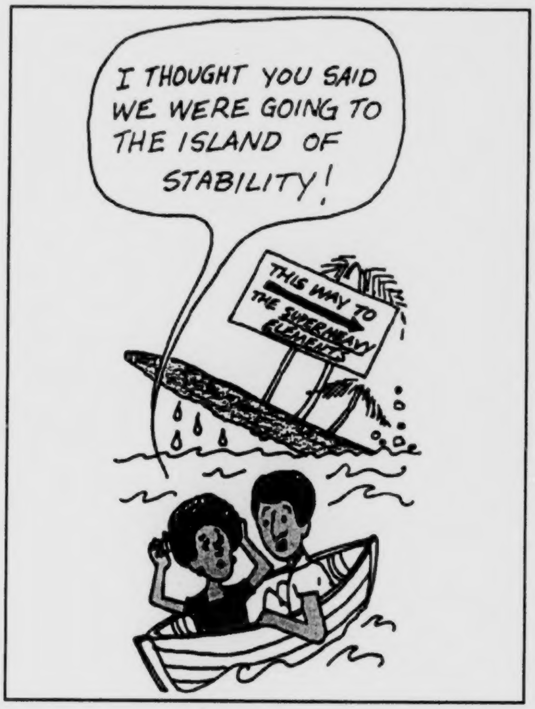
I THOUGHT YOU SAID WE WERE GOING TO THE ISLAND OF STABILITY!
THIS WAY TO THE SUPERHEAVY ELEMENTS

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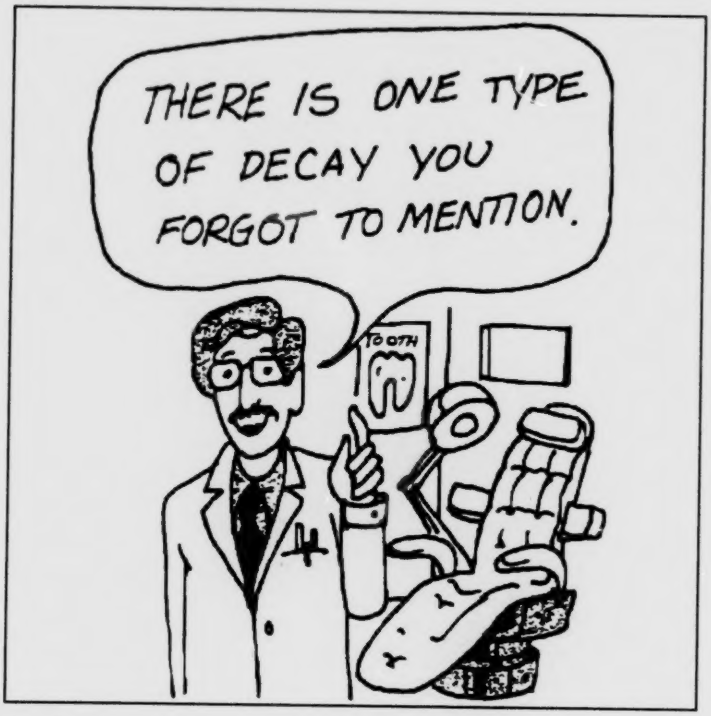
THERE IS ONE TYPE OF DECAY YOU FORGOT TO MENTION.
TOOTH

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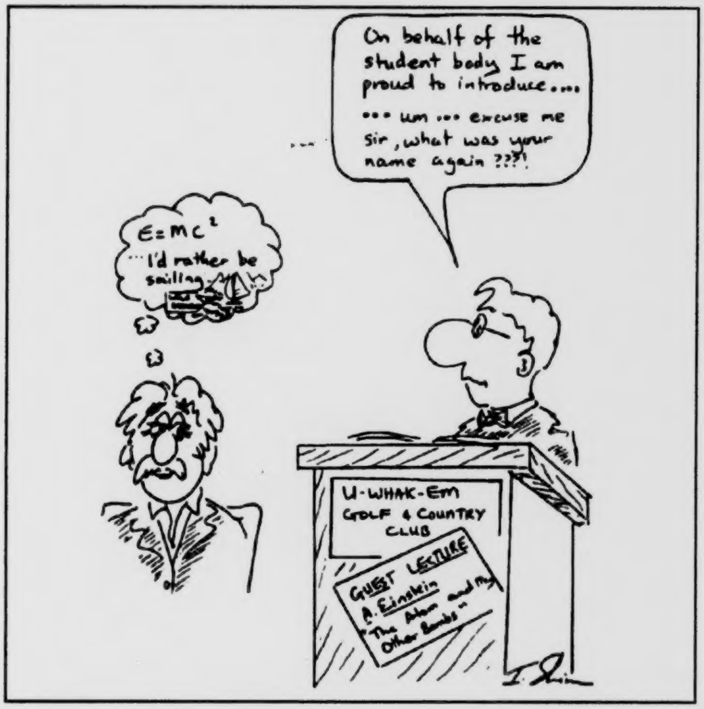
33

Shelly
"HE'S THE HALF-LIFE OF THE PARTY"

Used with permission of McGraw-Hill Book Co.

GEE, I CAN'T FIND A SEA OF INSTABILITY ANYWHERE ON THIS!

Used by permission of Veritex Publishing Co.


On behalf of the student body I am proud to introduce....
... um ... excuse me sir, what was your name again ???!
E=mc²
...I'd rather be sailing...
U-WHAK-Em GOLF & COUNTRY CLUB
GUEST LECTURE
A. Einstein
"The Atom and my Other Bombs"

CHEM 13 NEWS, September 1988, p. 47. Reprinted with permission.

A SourceBook Module

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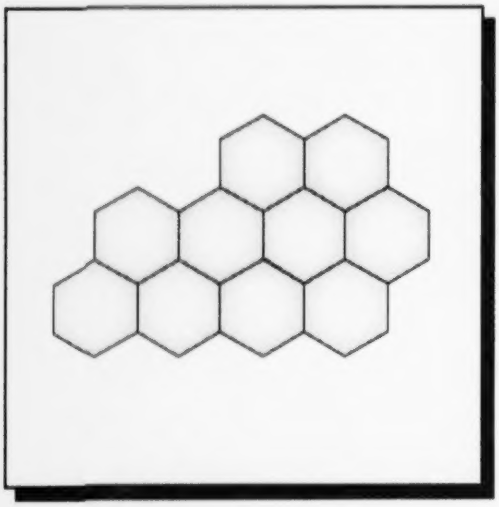
ORGANIC CHEMISTRY



ChemSource

Instructional Resources for Preservice and
Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

Organic chemistry is the study of carbon containing compounds. All living things are composed of carbon compounds. Until the early nineteenth century, many chemists believed that organic compounds must have their origin in living organisms, and consequently, could never be synthesized from inorganic material. Today, many compounds of carbon are still obtained from plant and animal sources. However, chemists have devised many processes by which natural and synthetic organic compounds can be prepared. Organic chemistry is a large and important field, and areas within organic chemistry are identified as disciplines themselves—biochemistry, polymer chemistry, synthetic organic chemistry, physical organic chemistry, *etc.* This module is an introduction to organic chemistry.

There are millions of known organic compounds, far more than the number of compounds not containing carbon because carbon atoms can form covalent bonds with one another (**catenation**) to an extent impossible for atoms of other elements. Carbon atoms participating in covalent bonding form four bonds for each carbon atom. Carbon atoms form branched and unbranched chains thousands of atoms long, and rings of different shapes and sizes. Carbon atoms form stable bonds with hydrogen, oxygen, nitrogen, the halogens, sulfur, and phosphorus. Many organic compounds are derived from petroleum.

Initially it might seem impossible to become knowledgeable about the chemistry of such a variety of compounds. Fortunately, there is a systematic approach to studying organic chemistry, allowing for the classification of organic compounds by the type of bond formed between adjacent carbon atoms and with the elements listed above. This systematic study classifies compounds according to "functional groups" contained in their molecular structure. Functional groups are portions of molecules (usually consisting of two to four atoms) responsible for characteristic physical and chemical properties of the substance. A list of names and structures for common functional groups in organic compounds is provided in Figure 1.

One important type of organic reaction is combustion (exhaustive oxidation or burning.) The products of combustion of organic compounds are carbon dioxide and water. The combustion reactions for methane (CH_4, natural gas) and butane (C_4H_{10}, used in Bic™ lighters) are:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$$

Combustion reactions are useful because they release a large amount of thermal energy. Almost all fuels used today are organic compounds. Other important groups of reactions are those which extend or combine carbon chains, or change a compound's functional group. Organic chemists have developed a wide variety of synthetic reactions used to build organic compounds and interconvert their functional groups. These reactions have enabled chemists to design molecules for specific purposes, as well as to systematically investigate the effect of structural changes on the properties of molecules.

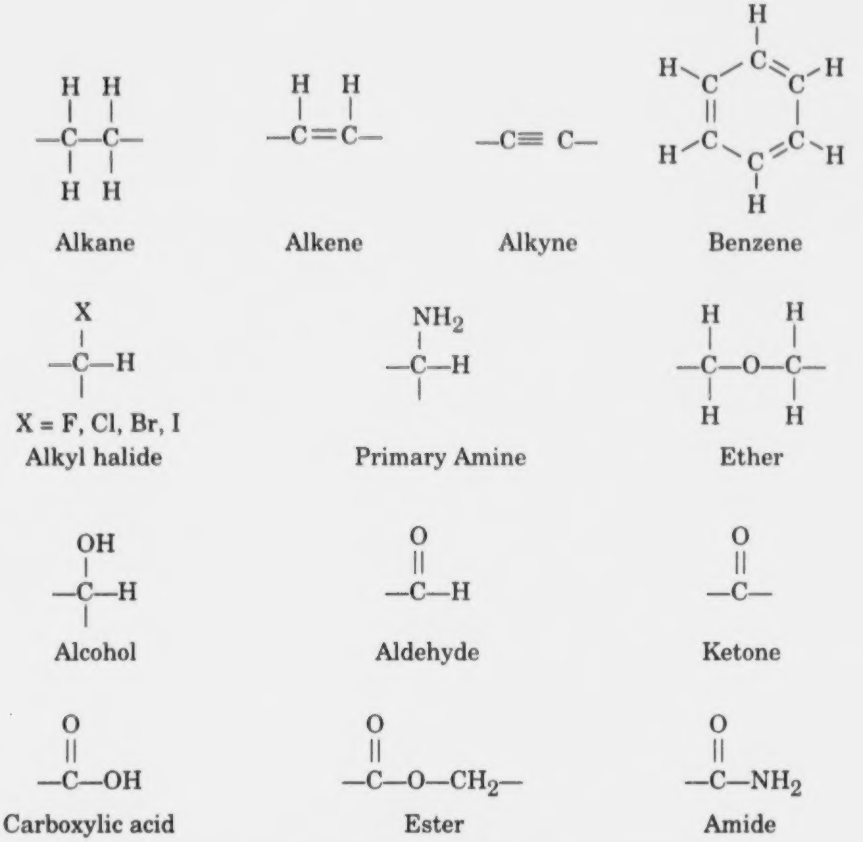


Figure 1. Common functional groups in organic compounds.

PLACE IN THE CURRICULUM

Traditionally, organic chemistry is not a major part of the first year chemistry curriculum. Most high school texts contain at least one chapter on organic chemistry, usually near the end of the book; however, there is no logical reason for this placement. To successfully study organic chemistry students will need as a prerequisite the conceptual understanding of the foundations of chemistry (atoms, molecules, stoichiometry, covalent bonding, phases of matter, *etc.*) The study of organic chemistry requires very little mathematical rigor. A good argument could be made for covering organic chemistry before topics such as acids and bases, oxidation-reduction, thermochemistry, kinetics, and electrochemistry.

CENTRAL CONCEPTS

1. Organic chemistry is concerned with carbon compounds containing covalent bonds. Since carbon atoms have four valence electrons, they form four covalent bonds.
2. The carbon-carbon bond can be a single, double or triple bond.
3. Functional groups are responsible for characteristic reactions of organic compounds.

4. Hydrocarbons are substances containing only carbon and hydrogen. They can exist as chains or rings.
5. Hydrocarbons react with oxygen in combustion reactions, producing energy, forming the basis for their use as fuels. Complete combustion forms carbon dioxide and water.
6. The synthesis of simple organic compounds can be accomplished by addition, substitution, and oxidation-reduction reactions.
7. Other important concepts are nomenclature, isomerism, stereochemistry, reaction mechanisms, and spectroscopy. These are important concepts but are not necessary or appropriate in an introductory course.

RELATED CONCEPTS

1. Atomic Structure
2. Relationship between electron structure and periodicity
3. Lewis-Dot structures
4. Brønsted-Lowry acid-base concept
5. Mole concept
6. Stoichiometry
7. Enthalpy
8. Equilibrium

RELATED SKILLS

1. Balancing equations.
2. Stoichiometric relationships.
3. Laboratory skills such as measuring mass and temperature and making solutions.

PERFORMANCE OBJECTIVES

After completing their study of organic chemistry, students should be able to:

1. define organic chemistry and give examples.
2. identify functional groups.
3. discuss the relationship of organic compound reactivity to functional group.
4. complete and balance hydrocarbon combustion reactions.
5. define the term "isomer" and give examples.
6. describe an organic synthesis that yields a useful product.
7. describe the preparation of esters from alcohols and carboxylic acids.
8. discuss ways in which organic compounds impact the environment.
9. give examples of organic compounds and reactions illustrating previously learned concepts.
10. exhibit safe laboratory practices for dealing with potentially flammable and toxic substances.

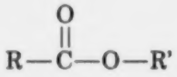
Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Synthesis of Esters

Introduction

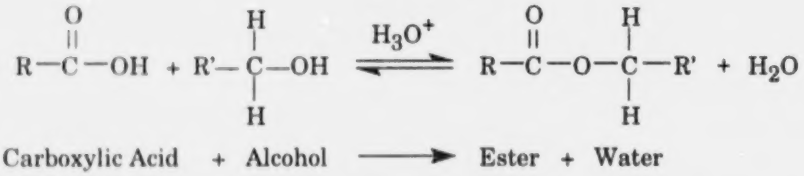
Esters are an important class of organic compounds which are characterized by the following generic formula.



$R-\overset{\overset{O}{\|}}{C}-O-R'$ R, R' = any alkyl group

Low molecular weight esters have pleasant odors and are responsible for many distinctive odors in fruits and flavorings. Esters can be readily prepared from a carboxylic acid and an alcohol. Generally this reaction is catalyzed by strong acids.

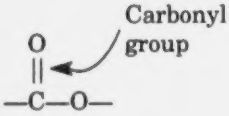
R, R' = any alkyl group



$$R-\overset{\overset{O}{\|}}{C}-OH + R'-\underset{H}{\overset{H}{C}}-OH \xrightleftharpoons{H_3O^+} R-\overset{\overset{O}{\|}}{C}-O-\underset{H}{\overset{H}{C}}-R' + H_2O$$

Carboxylic Acid + Alcohol ⟶ Ester + Water

The equilibrium constant for the reaction between primary alcohols and unhindered carboxylic acids is approximately 4. If equal quantities of acid and alcohol are used, the reaction gives a product yield of only 67%. In order to make esters in high yields, the equilibrium must be shifted toward products. According to LeChatelier's principle, an excess of alcohol or carboxylic acid, or removal of water will produce more ester. Sulfuric acid is used as a catalyst for this reaction because it also serves as a dehydrating agent, shifting the equilibrium toward products. The structures in Figure 2 will assist in determining which acid and alcohol will produce the ester with the desired flavor. The group to the left of and including the carbonyl group belonged to the original carboxylic acid, and the group to the right of the singly bonded oxygen, and including the oxygen atom, belonged to the original alcohol.



Purpose

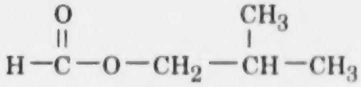
To synthesize an ester from the corresponding carboxylic acid and alcohol, and to identify the characteristic odor of each ester you have made.

Safety

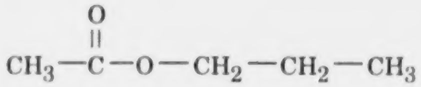
1. Wear protective goggles throughout the laboratory activity.
2. Be very careful with the concentrated sulfuric acid. It can cause serious chemical burns.
3. Your teacher will demonstrate the correct method for testing for odors using a wafting technique; you should never directly smell any chemical used or prepared in the laboratory.

Procedure

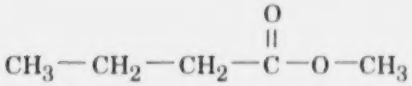
1. Place 75 mL water in a 100-mL beaker, add a boiling stone, and set the beaker on a wire screen on a ring stand. Light a burner and place it under the beaker. Bring the water to a gentle boil. (If available, hot plates should be used.)


$$H-\overset{O}{\overset{\|}{C}}-O-CH_2-\overset{CH_3}{\overset{|}{CH}}-CH_3$$

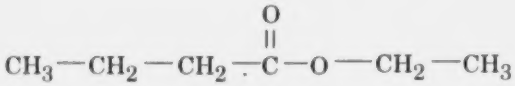
Raspberry


$$CH_3-\overset{O}{\overset{\|}{C}}-O-CH_2-CH_2-CH_3$$

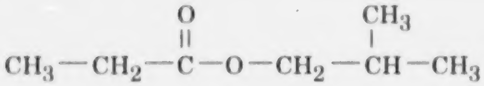
Pear


$$CH_3-CH_2-CH_2-\overset{O}{\overset{\|}{C}}-O-CH_3$$

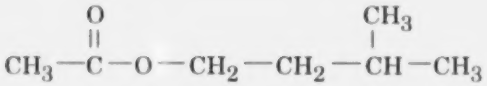
Apple


$$CH_3-CH_2-CH_2-\overset{O}{\overset{\|}{C}}-O-CH_2-CH_3$$

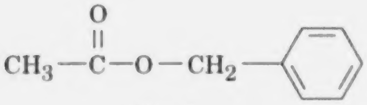
Pineapple

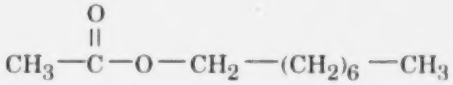

$$CH_3-CH_2-\overset{O}{\overset{\|}{C}}-O-CH_2-\overset{CH_3}{\overset{|}{CH}}-CH_3$$

Rum

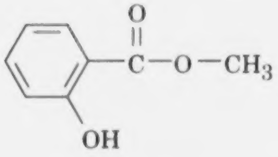

$$CH_3-\overset{O}{\overset{\|}{C}}-O-CH_2-CH_2-\overset{CH_3}{\overset{|}{CH}}-CH_3$$

Banana


Peach


$$CH_3-\overset{O}{\overset{\|}{C}}-O-CH_2-(CH_2)_6-CH_3$$

Orange


Wintergreen

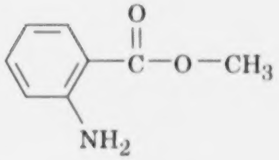

Grape

Figure 2. Esters. *NOTE:* —CH_2— is shorthand for $-\overset{H}{\overset{|}{\underset{H}{\underset{|}{C}}}}-$ *etc.* (See *Common Student Misconceptions*, p. 22, No. 4.)

2. Examine the structure of the ester you wish to synthesize. Determine which carboxylic acid and which alcohol you need to use in your synthesis.
3. In a 10-cm test-tube obtain one drop of carboxylic acid. If the carboxylic acid you wish to use is a solid, use about the amount that would fit inside this letter "O." Add three drops of alcohol and one drop of concentrated sulfuric acid. *Be very careful with the sulfuric acid. It can cause severe burns if spilled on the skin.*
4. Drop a boiling stone into the mixture.
5. Before heating, wave the fumes from this mixture to your nose and carefully note and record the odor of this reaction mixture.
6. Stopper the test-tube with a one-hole cork stopper and set the assembly in the boiling water. If the test-tube will not "stand up" in the beaker without getting water in the tube, put a test-tube clamp around the tube to hold it upright in the beaker.

7. Turn off the heat and allow mixture to stand in the hot water for 5-10 min.
8. Use a test-tube clamp to remove the reaction assembly from the water bath.
9. Fill a Beral™ polyethylene pipet with cold water and empty its contents into the test-tube. Mix the solution by gently swirling the test-tube.
10. Remove a small quantity of the aqueous mixture from the test-tube with the pipette. One way to do this without exposing yourself to the mixture in the tube is to push the tip of the pipette into the cork until the pipet shaft fills the hole in the stopper. Turn the tube so that the liquid in the tube covers the pipet tip and draw up some of the mixture. Then turn the test-tube right side up and withdraw the pipet.
11. Once the pipet is removed from the tube, invert it to let the liquid flow from the shaft to the bulb. Clear the liquid out of the shaft by lightly tapping the side or by drawing some air into the pipet.
12. With the pipet in the inverted position carefully force out some of the vapor and wave the fumes to your nose. Carefully note and record the odor of the ester you synthesized.
13. Occasionally, the vapor is too concentrated and it will "overpower" your sense of smell. Add more water to the mixture in the test-tube to dilute the odor.
14. If the odor is too faint, place the bulb of your pipette in warm water, wait for a minute, and try again.
15. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

Compare your ester's odor with the odor you expected from the name (grape, orange, *etc.*) Determine the ester's structure from the list above. Write a complete balanced equation for the synthesis of your ester.

Implications and Applications

1. In some cases the odor of the synthesized ester does not exactly match the odor of the fragrance found in nature. What might be a possible explanation?
2. Some of the compounds in the list of fragrances have functional groups in addition to the ester group. Identify these compounds and determine the types of functional groups present with the help of the functional group chart posted.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Synthesis of Esters

Major Chemical Concept

In this laboratory activity students will perform a standard organic synthesis—producing esters from alcohols and carboxylic acids. By examining the chemical structure of the desired ester they will choose the appropriate starting materials for their synthesis. They will analyze their reaction mixture for presence of the ester through its characteristic odor.

Level

Basic, General, or Honors

Expected Student Background

Students should be able to identify the alcohol, carboxylic acid and ester functional groups. They should be familiar with the basic bonding characteristics of organic compounds and understand the concept of chemical synthesis. An understanding of catalysis, equilibria and LeChatelier's principle is desirable but probably not required.

Time

Approximately 30 min.

Safety

Read the *Safety Considerations* in the *Student Version*. A major concern with this laboratory is asking students to "sniff" their products. Because of the small quantities of reagents used in this procedure, this hazard is minimized. Try the technique, described below in the *Pre-Laboratory Discussion* section. It is important to clear the barrel of the pipet before smelling the vapors from the ester solution. *It is mandatory that you demonstrate the proper technique for detecting odors before the laboratory period begins.* All organic compounds used in this laboratory are flammable. It is imperative that reagents be away from heat sources and students use small quantities to minimize fire danger. It is preferable to use hot plates rather than burners.

Materials (For 24 students working in pairs)

12 Test-tubes, 10-cm, fitted with one-hole *cork* stoppers (do not use rubber stoppers)
12 Large-barrel polyethylene pipets (*e.g.*, Beral™)
12 Beakers, 100-mL
12 Ringstands, wire screens and burners or hot plates
Boiling stones
12 Labeled 10-mL bottles of concentrated sulfuric acid (H_2SO_4), with polyethylene pipet
6 Labeled test-tubes, 10-cm with pipets, containing the following alcohols: methanol, ethanol, 1-octanol, 2-methyl-1-butanol, 2-methyl-1-propanol, 1-propanol, 3-methyl-1-butanol, benzyl alcohol.
6 Labeled test-tubes, 10-cm with pipets, containing the following acids: formic acid, acetic acid, propionic acid, butyric acid, salicylic acid, anthranilic acid.

NOTE: Structural formulas should be on the carboxylic acid and alcohol labels. Butyric acid has a very foul odor and should be used with caution and very good ventilation.

Advance Preparation

Set the reagents in a fume hood to keep the background room odor to a minimum. If possible, open your windows to help circulate the room air. It is not necessary to provide the reagents for all of the esters to be synthesized. A subset of three of four esters can be used (*e.g.*, pineapple, banana, wintergreen, orange.) Label both reagent containers *and* dispensers to avoid contamination. A good way to keep the reagent container and dispenser together is to tape a test-tube to the reagent bottle and place the polyethylene pipet used to dispense the reagent in the test-tube. The cork stoppers that are used in this laboratory need to accept the polyethylene pipet with a tight fit. Polyethylene pipets with tapered ends work nicely. You probably will have to dispose of corks after each laboratory because they pick up the odor of the synthesized ester and will confuse the next set of students if recycled. If your cork borer does not drill a small hole, you can clamp the solid cork stoppers in a vise and drill holes in them with an electric drill fitted with a small drill bit. Have a copy of Figure 1 posted to identify functional groups.

Pre-Laboratory Discussion

It is important to demonstrate the proper method for detecting odors. In one hand hold the open test-tube about 15 cm from your nose. Using your free hand sweep the air space above the test-tube toward your nose.

Discuss the concept of organic synthesis, particularly the synthesis of esters (called esterification). The mechanism for ester synthesis is well established but discussing this mechanism is not appropriate for first year students. You might have students try to determine the systematic name for the esters which they make. This would help determine which carboxylic acid and alcohol to use for the synthesis. You will notice that some compounds have common names. To avoid confusion be careful in labeling the compounds.

For basic or general chemistry students, tell them which carboxylic acid to mix with which alcohol to produce a particular ester. Alternatively support student learning by having students predict in the pre-laboratory which ester is produced by particular combinations of carboxylic acid and alcohol, and to compare the structure to the table of esters. Once students have predicted the structure and odor of particular combinations, assign combinations to student laboratory groups to synthesize.

Teacher-Student Interaction

Help students observe physical and chemical changes that take place when the reaction mixture is being heated. Discuss the purpose of the boiling stone *[provides a site for bubbles to form during boiling]*. Discuss the concepts of catalysis, equilibrium and LeChatelier's principle.

Anticipated Student Results

Students should recognize odors like banana and wintergreen. Other odors may be less readily identified. (Undercooking may be better than overcooking.) Students' odor palates are easily overwhelmed and may confuse odors. Students should realize that natural odors are mixtures. Many students will be able to write the reaction equation for any ester after writing the equation for their ester; others will need to see more than one example before they can generalize.

Answers to Implications and Applications

1. The fragrances found in nature are most often due to a mixture of esters rather than a single compound. Fresh ground coffee has at least 200 identifiable esters.
2. –OH hydroxyl, $-NH_2$ amine, $-C_6H_5$ phenyl group (aromatic compound)

Post-Laboratory Discussion

This activity is designed to provide students with hands-on experience in conducting an organic synthesis. Students should be able to discuss the industrial role of organic synthesis in making compounds which are difficult or expensive to isolate from naturally occurring sources. Vanilla is an example of a synthesized flavoring. Aspirin is a familiar example of a synthesized drug (see *Chemistry in Medicine* module). Most of the vanilla and all the aspirin used today are synthesized. Students should also be asked to discuss the benefits and hazards of preparing synthetic compounds not found in nature. Various pesticides, herbicides, and polymeric substances are everyday examples. Discuss benefits and abuses of these substances.

Extensions

1. The ester functional group is very important in the synthesis of polymers. Look up the chemical structure of polyester polymers (see *Polymers* module) and determine the compounds used in their synthesis.
2. Look on the label of common household products to see if you can identify the esters.
3. Suppose you looked on a bottle of fingernail polish remover and found it was composed of ethyl acetate. How would you know it was an ester? How would you prepare this ester? *[Smell. Can be prepared from ethanol and acetic acid.]*
4. Make molecular models of a carboxylic acid, an alcohol, and the ester that can be synthesized from them.

Assessing Laboratory Learning

1. Have students write equations for the formation of esters not synthesized during their laboratory period.

Raspberry	2-methyl-1-propyl formate
Pear	1-propyl acetate
Apple	Methyl butyrate
Pineapple	Ethyl butyrate
Rum	2-methyl-1-propyl propionate
Banana	2-methyl-1- butyl acetate
Peach	Benzyl acetate
Orange	Octyl acetate
Wintergreen	Methyl salicylate
Grape	Methyl anthranilate

2. Have students look up the chemical structure of aspirin (an ester) and write an equation for its synthesis.

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: Chemical Analysis for the Presence of Halogens

Introduction

Halogens can be substituted for hydrogen in organic molecules to form compounds possessing physical and chemical properties that are different from the original unsubstituted compounds. Since halogens have higher electronegativities than hydrogen, these substituted compounds are frequently polar in nature. These "halogenated hydrocarbons," or alkyl halides, are used as reagents in several important organic synthesis reactions. Polyhalogenated alkyl compounds (those containing more than one halogen) have practical applications as dry cleaning solvents, insecticides and fire extinguishers. Until recently, chlorofluorocarbons or CFCs have been used as propellants in aerosol spray cans. Unfortunately, these propellants contribute to the decline of ozone in the earth's stratosphere. Stratospheric ozone absorbs harmful ultraviolet light that is emitted by the sun. The decrease of stratospheric ozone is a critical environmental issue.

Presence of chlorine, bromine and iodine in organic compounds is detected by the Beilstein test. A positive Beilstein test results from the formation of a volatile copper halide when an organic halide is heated with copper oxide. The copper halide imparts a blue-green color to a flame. Since the reaction occurs at a high temperature in a flame, the exact products have not been identified. The Beilstein test does not differentiate between chlorine, bromine and iodine. An additional test using sulfuric acid and sodium hypochlorite can be used to identify which of these halogens is present.

Purpose

To test several unknown organic compounds for the presence of a halogen using the Beilstein test and to identify the halogen.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Your teacher will demonstrate the proper technique for detecting odors of volatile chemical compounds.
3. Many of the organic compounds used as unknowns are flammable. Since this laboratory uses burners care must be exercised to avoid fires.
4. Care should be exercised in handling the sulfuric acid and sodium hypochlorite solutions.

Procedure

1. Obtain from your instructor an unknown sample of an organic compound. Record the number which identifies the unknown in your laboratory notebook.
2. After bending a small loop (about 2 mm in diameter) in one end of a 5-10 cm length copper wire, insert the other end into a cork stopper (to be used as a holder).
3. Holding the cork, heat the loop end of the wire in a burner flame.
4. After cooling (approximately one minute), dip the wire directly into the sample of the unknown substance.
5. Heat the wire in the burner flame again. The compound will first burn, probably with a yellowish flame. After burning, a green flame will be produced if a halogen is present.

6. If a halogen is not present, record your results and obtain another unknown from your instructor. If a halogen is present, proceed to Step 7.
7. Place 2 mL 1 M sulfuric acid in a 20-cm test-tube.
8. Add 5 drops of your unknown organic halide (for solids, use size of a matchhead) and a boiling stone to the test-tube.
9. Heat this mixture for 5 min in a boiling water bath and allow the test-tube and its contents to cool in a beaker of cold water.
10. Add 10 drops Trichlorotrifluoroethane (TTE) and 3 drops sodium hypochlorite solution to the test-tube.
11. Check to see that the solution is still acidic with a piece of moist blue litmus paper. If it is not acidic, add dropwise 1 M sulfuric acid to make the solution acidic.
12. Stopper the solution and shake it vigorously. Set it aside and allow the layers to separate. Separation should occur within 10 min.
13. Record your observations and interpretation in your laboratory notebook. (The TTE layer will dissolve the colored halogens. An orange to brown color in the TTE layer indicates the presence of bromine. A violet color indicates the presence of iodine. No color or a light yellow indicates the presence of chlorine.)
14. Dispose of your solution in the container marked "halogenated waste."
15. Thoroughly wash your hands before leaving the laboratory.

Data Analysis

Perform this test on several different unknown organic liquids provided by your teacher. Report your results to your instructor. During the next class period your instructor will tabulate the class results. Compare your results with those of others.

Implications and Applications

Attempt to identify halogenated compounds which are in common use today. Read product labels in grocery, garden and hardware stores to find products which contain halogenated organic compounds. Discuss the use of these products, their toxicity, and the proper means of disposing of them.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: Chemical Analysis for the Presence of Halogens

Major Chemical Concepts

This activity is designed to allow students to test for the presence of halogens in organic compounds. The student will be required to make careful observations and to compare these observations for several unknowns. Students will then infer the absence or presence and identity of halogens in their samples.

Level

This laboratory exercise is appropriate for general students late in the academic year or for second year chemistry students.

Expected Student Background

Students should have standard laboratory skills and a general understanding of the structure of organic compounds. You should discuss halogenated hydrocarbons and their properties in class before this laboratory is performed.

Time

40 min

Safety

Read the *Safety Considerations* in the *Student Version*. Check the material safety data sheets (MSDS) or ChemSource *Safety* module, Appendix II for all of the unknowns used in this laboratory exercise. (Many halogenated hydrocarbons are hazardous compounds and should not be used as unknowns. The substances suggested in this guide are nontoxic and noncarcinogenic.) Use small quantities of unknowns to avoid any hazard which could result from spillage. If possible, store all reagents in a fume hood and allow students to take only the amount of sample needed for each activity.

Materials (For 24 students working in pairs)

12 Burners
12 Forceps
36 Test-tubes (15-cm length)
12 Test-tube holders
12 400-mL Beakers
12 Wire gauze
12 Ringstands and rings
Boiling stones
12 Corks
12 Graduated cylinders, 5- or 10-mL size
Copper wire, 12 pieces, 7-8 cm
36 Pipets with bulbs
1 M Sulfuric acid, H_2SO_4, 50 mL (5 mL conc. H_2SO_4 diluted to 50 mL)
1,1,2-Trichloro-1,2,2-trifluoroethane (TTE), 20 mL
Sodium hypochlorite solution, NaClO (use Clorox™ or comparable bleach), 20 mL
Unknown Solutions (see *Advanced Preparation*)
Blue litmus paper
Waste bottle for unknown mixtures

Advance Preparation

Make up unknown solutions as 10% solutions of organic compounds in appropriate solvents (cyclohexane for nonpolar substances and acetone for polar substances). Have sufficient unknowns for students to perform tests on three to five different compounds (1 g/10 mL is enough for 24 students. Use 5-10 drops of unknown per test). Suggested unknowns for halogens are 2-iodobenzoic acid, 2-bromobenzoic acid and 4-chlorobenzoic acid. (These are commonly used as preservatives and can be obtained from Aldrich Chemical Company, 1001 West Saint Paul Avenue, Milwaukee, WI 53233, as well as other suppliers.) Over-the-counter pharmaceutical products can also be used. Examples which are antihistamines include chlorpheniramine (found in Dristan™) which contains chlorine; brompheniramine (found in Dimetane™) which contains bromine; and clioquinol (found in the antifungal product, Vioform™) which contains both iodine and chlorine. Other household products can be tested from paint or hardware stores. Examples include cleaning agents, solvents, white-out, *etc.* Suggested unknowns for nonhalogens are alcohols, esters, and benzoic acid. Be sure to clearly label the unknown containers with A, B, C, D, *etc.* It is advisable to prepare a data table for the students. *NOTE*: Larger amounts of bromide are needed to give a good observable test.

Pre-Laboratory Discussion

Discuss with the class the different functional groups of organic compounds. Explain that halogens form single bonds with carbon atoms in organic molecules.

Teacher-Student Interaction

One approach to teaching this laboratory is to demonstrate the laboratory using two known substances. One known (*e.g.*, benzoic acid) should not contain a halogen, and a second known (*e.g.*, bromobenzoic acid) should give a distinctive positive test. Students can then observe definitive negative and positive tests, respectively. Allow students to perform tests on the known samples as part of their laboratory exercise.

Anticipated Student Results

1. If students use a clean copper wire for each test, a positive or negative halogen test should be apparent.
2. In distinguishing among halides, the colors of I_2 and Br_2 are less apparent. Students may need to hold their TTE extraction tubes against a piece of white paper that is backlighted from a window or a lamp in the classroom.
3. Students will enjoy comparing their analyses of household items such as mothflakes with actual label information.

Answers to Data Analysis

The answers will depend on the type of halogens used.

Post-Laboratory Discussion

1. Help the class compile a list of halogenated organic compounds in common use today. Freon, a refrigerant, is an example.
2. Discuss the use of pesticides in the agriculture industry and the impact of various halogenated compounds on the environment (*e.g.*, polychlorinated biphenyls (PCB's), chlorofluorocarbons—CFCs, halon fire extinguishers, DDT, *etc.*).

Extensions

1. Discuss the principles involved in the reactions used to identify the presence of halogens (conversion of organic halide to inorganic [copper] halide) and the identity of the halogen (conversion of organic halide to inorganic halide and the subsequent oxidation of halide ion to the free halogen by ClO^-). Reactions for the Beilstein test are not well understood; however, the reactions for the identification of the halogen could be researched. The oxidation of halide ion by hypochlorite ion provides an opportunity to review redox reactions.
2. Locate the material safety data sheets (MSDS) that are shipped with chemical orders. Share them with your class and compare the data for several different classes of organic compounds including halogenated compounds. Take special notice of the proper disposal requirements for these substances.
3. The number of double bonds in an unsaturated hydrocarbon can be determined by using the "iodine number." Investigate how the "iodine number" can be used to determine whether oils and lard have single or double bonds.
4. At one time, carbon tetrachloride was used widely as a cleaning agent. It is now considered to be a carcinogenic compound. What halogenated hydrocarbon replaced carbon tetrachloride in the cleaning industry? (*HINT: Have students visit a dry cleaning establishment.*)

Assessing Laboratory Learning

This activity will allow students to develop increased confidence in their powers of observation. They should be capable of drawing conclusions from observations for several different results. Most of this laboratory exercise is designed to teach science process skills. These skills can be assessed only by observing students as they perform the various tasks. Ask students to describe their observations and compare the observations of several students for each unknown tested. It is important that students be allowed to repeat their tests on many different unknowns. Both negative and positive Beilstein results should be obtained in order to allow students to differentiate between them. Encourage students to perform a test together whenever there is disagreement regarding a particular observation. The emphasis should not be on "right answers" in this activity. Rather than post the "answers" to the unknown solutions, see if the class can agree upon the results of their experimentation.

DEMONSTRATIONS

CAUTION: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Methane Bubbles

Purpose

This demonstration shows the reactivity of saturated hydrocarbons.

Materials

Methane gas (natural gas) outlet
Glass funnel, 5-7 cm diameter, with 1 m length of rubber tube
Glycerol, 10 mL, mixed with 10 mL liquid detergent in a 100-mL beaker
Candle attached to the end of a meter stick
Matches
Safety goggles

Safety

CAUTION: Methane is a flammable gas. Make sure students stand back and that the gas jet is turned off before igniting the bubble. Designate one student as the fire control officer with the fire extinguisher readily available.

Procedure

Connect the tubing to a natural gas outlet with the other end attached to the end of a glass funnel. Dip the funnel in the detergent/glycerol mixture and form a bubble by opening the gas tap slightly. Turn off the gas. Turn the funnel right-side up and shake it gently to allow the bubble to rise. Light the candle and touch it to the bubble. Observe the reaction and ask students to name the products. *NOTE*: An alternate procedure is to bubble methane directly through detergent solution to form a "tower of bubbles" that can be scooped aside and ignited.

Remarks

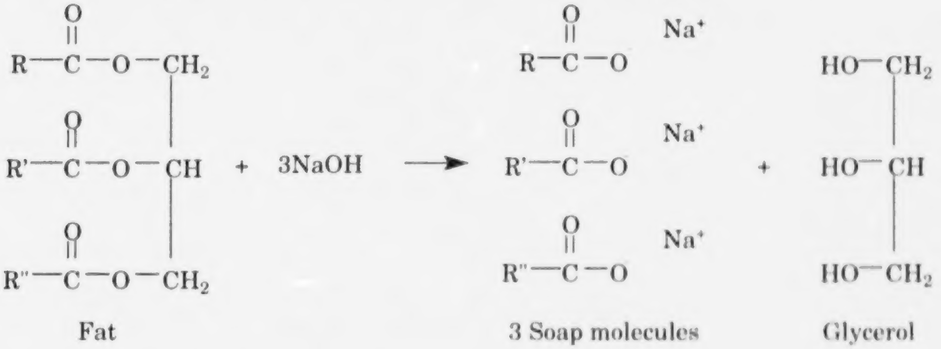
Glycerol lends structural stability to the bubble.

Demonstration 2: Making Soap

Introduction

This demonstration will require one to three days for completion, but you should be able to conduct other activities during these class periods.

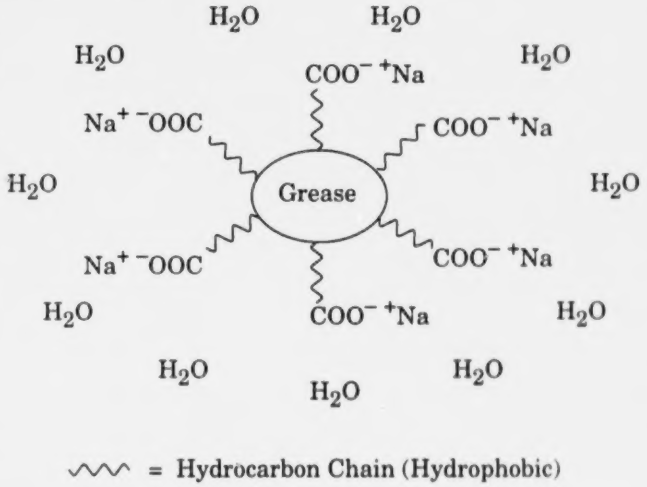
Soap has been known at least since 600 B.C. when the Phoenicians reported making a curdy material by boiling animal fat with wood ashes. Soap is prepared by allowing sodium or potassium hydroxide to react with fats and oils (esters of long chain carboxylic acids with glycerol).



(R, R', R'' represent long-chain hydrocarbons)

The soap curds produced contain both glycerol and excess base. Failure to remove the base makes homemade soap harsh. However, soap can be purified by boiling it with plenty of water, followed by precipitating the dissolved soap with sodium chloride. (The process of precipitating a substance by dissolving another substance that is more soluble is called "salting out.") The precipitated soap can be dried and pressed into cakes. Perfumes and dyes are usually added to soap intended for household use; antiseptics are added to make medicinal soap; pumice is added to make scouring soap; and air bubbles are blown into soap to make floating soap. Despite advertising claims, soaps are quite similar; the main difference is the kind of fat or oil used in their synthesis.

Soap works as a cleanser because the ends of soap molecules have different properties. The carboxylic acid salt, or carboxylate, end is ionic and dissolves in water, a property described as hydrophilic or water-loving. The long hydrocarbon chain part of the molecule is not soluble in water (hydrophobic), but it is soluble in other nonpolar substances such as fats and greases. Even though the greasy dirt won't dissolve in water, the addition of soap provides a mechanism for water to carry greasy dirt away. Each little glob of grease is surrounded by soap molecules which have their hydrocarbon chains dissolved in the grease glob. Water then surrounds and dissolves the carboxylate ends of the soap molecules which are sticking out on the surface of the grease glob. The net effect is that the grease glob is now soluble in water.



Commercial cleansers, are made of synthetic detergents rather than natural soaps. Detergents are not the same as soaps derived from fats, although they are frequently referred to as "soaps." Modern detergents are typically mixtures of compounds called surfactants that have a hydrophilic "head" and a hydrophobic "tail." Detergents typically contain polyphosphate groups that serve to soften hard water and to adjust the pH. Laundry detergents also include fluorescent "whiteners" that absorb ultraviolet light and then emit visible light to make clothing appear cleaner.

Purpose

This activity demonstrates the preparation of soap using ordinary household materials.

Materials

Oil, shortening, or lard, 20 mL (a mixture of lard and olive oil works quite well)
6 M Sodium hydroxide, NaOH. (To a 125-mL or 250-mL pyrex flask add 24 g NaOH and 100 mL H_2O. Dissolve the NaOH by stirring. Be careful. Dissolving NaOH liberates heat.)
Ethanol, C_2H_5OH
2 Beakers, 100-mL
Thermometer
Hot plate *(Do not use an open flame for this activity!)*
Wooden or plastic spoon
Paper cup
25% Sodium chloride, NaCl (25 g NaCl in 75 mL H_2O)
Beaker, 600-mL
Beaker, 1000-mL (optional)
Sodium chloride, NaCl, 50 g (optional)
Perfume (optional)
Food coloring (optional)
Ice bucket and ice (optional)
pH paper (optional)

Safety

Ethanol is flammable. Do not use open flames. Sodium hydroxide is caustic; protect your skin from splashes and wash off immediately if a splash occurs. Unpurified soap is often too basic for human use. Test the product with pH paper. If the pH is above 8, don't let students wash with it. Wear an apron, goggles, and gloves. If you have a face shield, wear that instead of goggles.

Procedure

1. In a 100-mL beaker, prepare about 20 mL fat mixture by melting lard or shortening. If liquid oil is used, measure out 20 mL in a graduated cylinder.
2. Allow the heated mixture to cool to about 45 °C.
3. While the liquid fat is cooling, heat a mixture of 10 mL 6 M NaOH and 10 mL ethanol to about 35 °C in a separate 100-mL beaker. Use care; do not splatter this solution—6 M NaOH is quite caustic. While the liquid fat is cooling and the ethanol and NaOH are heating, pour the aqueous NaCl into the 600-mL beaker.
4. When the ingredients are at the desired temperature, slowly pour the ethanol solution of NaOH into the melted fat while stirring continuously with a wooden or plastic spoon.
5. Stir constantly until a thick substance in which the spoon will stand upright is produced. If the reactants were at the right temperature, this will take 5-10 min. [If soap does not form, heat the solution in a water bath made by heating water in a 250 or 400-mL beaker and placing the 100-mL reaction beaker in it. Heat the reaction in simmering water for about 20 min. Then pour the warm reaction mixture into the 25% NaCl solution that was previously put in the 600-mL beaker.] As the mixture cools, the hardened soap will float to the top where it can be skimmed off and placed in a separate container. If soap forms readily, omit the step with NaCl. Go to either Step 6 or 10 next. Step 10 begins the purification procedure.
6. Add perfume and color at this point. Just stir it in. (Optional)
7. Pour the mixture slowly and evenly into the paper cup—the soap mold.
8. Test the sudsing power of prepared soap by washing out the beaker in which it was prepared.
9. Let the soap settle for a few days and dry into the shape of the cup.
10. To purify soap, heat to 50-80 °C approximately 500 mL of water in the 1000-mL beaker. Then pour the soap in to dissolve it. Let it cool.
11. Add 50 g NaCl and stir until dissolved. The soap should precipitate out. To increase the amount and rate of soap precipitation, cool the solution in an ice bucket.
12. Pour off excess water, stir in perfume and color, and proceed as in Steps 7-9.
13. If time permits, prepare two samples of soap, one with and one without purification. Use pH paper to test the pH of both samples. This is a good way to determine whether soap is still too basic for washing skin. Avoid using soap at pH > 8.

Questions

1. Describe the reaction in terms of reactants and products. What functional groups are involved? *[An ester of glycerol and carboxylic acids with long alkyl chains reacts with sodium hydroxide to produce glycerol and the sodium salt of the carboxylic acid. The ester functional group reacts with the hydroxide ion.]*
2. What was the pH of soap? Why? How could it be purified further to alter the pH? *[The pH of the soap will vary. pH is high because of sodium hydroxide impurity. It can be lowered by dissolving the soap in water and re-precipitating it by salting out. The NaOH contaminant remains dissolved in the water.]*
3. A popular drain cleaner is a concentrated solution of sodium hydroxide. How do you think it works to remove grease from the drain? *[The sodium hydroxide in drain cleaner reacts with grease in the drain to turn it into soap. The soap formed dissolves in water and washes down the drain.]*

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. What is the role of carbon-carbon bonding in organic chemistry? *[Provides for a variety of organic compounds (open-chain and cyclic) with a multitude of uses.]*
2. What is the role of functional groups in organic chemistry? *[Provides the major route for converting one organic compound into another.]*
3. What are some reactions of organic compounds that lead to new products? *[Substitution reactions lead to the formation of artificial flavors (*Activity 4*) and soaps (*Demonstration 2*).]*
4. What role does combustion of organic compounds play in modern society? *[Combustion reactions of organic compounds is exothermic providing large quantities of heat. This heat can be used to warm buildings, boil water, run automobile engines,* etc.*]*
5. How are structure of an organic compound and its physical and chemical properties related to each other? *[The structure of an organic compound and its physical and chemical properties are related in terms of water solubility or insolubility, density, physical state, and reactivity.]*

Counterintuitive Examples

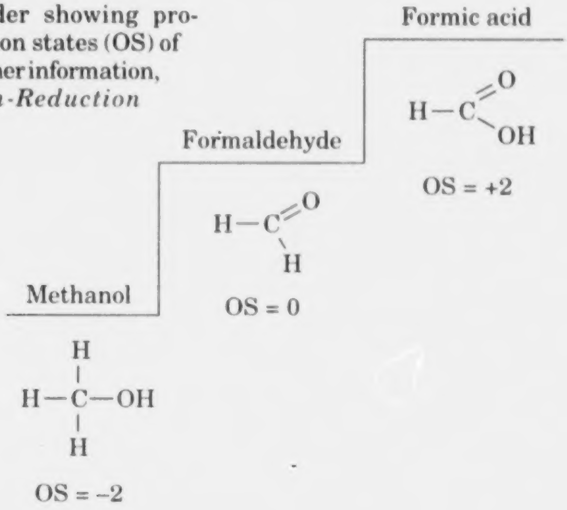
1. Small changes in structure produce large changes in perceived properties, such as odor. For example, esters have small structural differences yet quite varied odors. Compare the structures responsible for the grape and wintergreen odors in *Activity 1*.
2. That all organic molecules come from living sources is a common (and incorrect) belief. However, as a chemical term, *organic* just means having a carbon-chain backbone, which can be synthesized from *inorganic* substances.
3. Most of what is taught in first-year chemistry is inorganic chemistry. However, most of the chemistry around us—fuels, dyes, detergents, plastics, biological processes—is organic chemistry.

Analogies and Metaphors

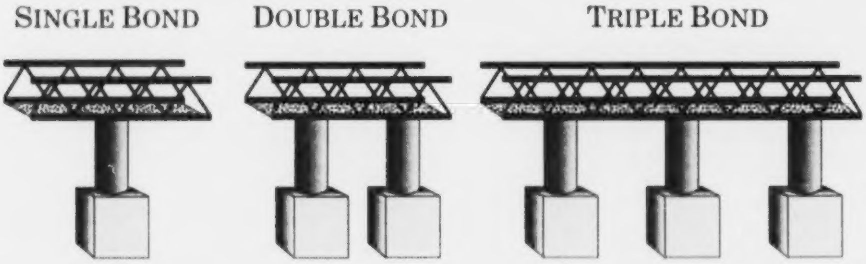
1. Some types of carbon compounds are like a linked chain—all connected together.
2. Organic synthesis is like building with LEGOs™ or CAPSELA™ kits. A few basic components are used to produce many end-structures.

Pictures in the Mind

1. Oxidation ladder showing progressive oxidation states (OS) of carbon. For further information, see *Oxidation-Reduction* module.



2. Bond strength and bridges. Remember that bond *strength* refers to how strongly two atoms are held together, *not* to how *reactive* those bonds are.



One beam supporting bridge

Two beams supporting bridge

Three beams supporting bridge

TIPS FOR THE TEACHER

Language of Chemistry

Much emphasis is placed on naming organic compounds in most introductory chemistry texts. Such emphasis is misplaced, and the emphasis in introductory chemistry should be determining chemical and physical properties through the presence of functional groups. Ask students to learn to recognize and name a few important functional groups such as alcohols, aldehydes, carboxylic acids, *etc.* The focus should then be on what these functional groups will do chemically because of the electron distribution in the group. Rather than focusing on language in the organic chemistry module, use organic compounds and reactions to illustrate the fundamental principles of chemistry.

An important term for students to learn is *isomer* because isomerism is very important in determining a compound's physical and chemical properties. Isomers that differ in identity of functional groups are clearly different from each other chemically and physically. However, even if the functional groups remain the same, as in *cis-trans* isomerism, a small change in geometry sometimes greatly alters the overall shape of the molecule, the distribution of electrons within the molecule, and/or the location of functional groups in space with respect to each other. Therefore, even *cis-trans* isomers can differ profoundly from each other. A number of isomeric puns are provided in the *Humor* section to illustrate some of the language of isomerism. Isomerism is also discussed in Links and Connections to the Contemporary World section.

Distilled alcoholic beverages (ethyl alcohol) are often described by their "proof." Since this term may be familiar to students, the history of this term is outlined. "Proof" refers to a test used to determine the alcohol content of whiskey. To "prove" that whiskey was not watered down, gunpowder and the whiskey in question would be mixed to form a paste, usually in a shot glass or even just in a pile on the bar. A lighted match would be placed on this paste. If the whiskey was much less than 50% alcohol, the match would just go out. If the whiskey was much more than 50% alcohol, the mixture would explode. (Whether this whiskey was considered good or bad to drink is not recorded in folklore.) If the whiskey was 50% alcohol and 50% water, the mixture would burn steadily. Such whiskey was usually called 100 proof. Later, when it became easy to determine the percentage of alcohol in an aqueous alcohol solution, it was discovered that 100 proof whiskey was 50% alcohol. ("Proof" is twice the percent alcohol.)

Pattern Recognition

Properties and biological activity are related to functional groups. To emphasize the functional group nature of organic compounds, make overheads of line structures and/or models of complex organic molecules like the porphyrin center of hemoglobin, or a steroid, or quinine, *etc*. You can find structures for such interesting molecules in college organic chemistry texts. Have students identify the various functional groups and predict chemical properties from them. For example, they can predict that quinine will be basic because of the amine groups with unshared electron pairs that can act as Lewis or Brønsted-Lowry bases (see *Acids and Bases* module). Given that it is predicted to be a base, you can then give students some quinine water in paper cups to taste and discuss the name "bitters," which is sometimes given to quinine water.

Students can identify the Lewis base amine centers in hemoglobin and show the complex with iron is formed by donating pairs of electrons to bond with iron.

The number of biologically important and/or interesting molecules that can be used to play this game are practically infinite, and your ability to use this exercise with students will be limited only by your own research and imagination.

Common Student Misconceptions

1. **"'Organic' means 'natural' or 'good'."**

   The word "organic" is used in many different contexts and with many different connotations. Advertisers sometimes take advantage of this confusion. For example, in the early 1970s, during the beginning of the environmental movement, one shampoo manufacturer advertised that their product was made from "pure one hundred percent organic chemicals." Of course, some of the most potent carcinogens known are "pure organic chemicals." As a chemical term, *organic* simply means having a carbon chain backbone.

2. **"All carbon containing compounds are organic."**

   Carbonates, bicarbonates, carbon dioxide and carbon monoxide are not organic compounds. Some inorganic compounds can produce an organic molecule, as indicated in the following reaction in which calcium carbide (inorganic) is hydrolyzed to acetylene (organic).

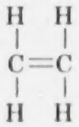
   $$CaC_2(s) + H_2O(l) \rightarrow C_2H_2(g) + Ca(OH)_2(s)$$

3. **"Formalin and formaldehyde are the same."**

   Formaldehyde is a gas at standard temperature and pressure. It is a very reactive molecule and is a carcinogen. However, *formalin* is an aqueous solution of formaldehyde. Because formaldehyde is so reactive, it reacts with the water to form a molecule that is not carcinogenic. The molecule in formalin solutions is so stable that formaldehyde is not present in its vapor. Formalin in high concentrations is used as a preservative for biological specimens. Although the odor of formalin is unpleasant, current evidence indicates that formalin is not a health hazard.

4. **"When a formula is written in one line such as $CH_2=CH_2$, the hydrogen atoms are bonded to the carbon atom to their left."**

   This shorthand notation may be misleading, but it is essential to be able to read it properly. The carbon on the left is bonded to two hydrogens. Since carbon has four covalent bonds, the double bond to the immediate right of the two hydrogens must apply to the left hand carbon as well. Furthermore, since each hydrogen only has one bond, then the hydrogens can't possibly be bonded to any other atoms but the carbon. The right side of the molecule is less ambiguous. The double bond applies to the right hand carbon, which must also be bonded to the two hydrogens on the right. An expanded version of this formula can be drawn out as shown.

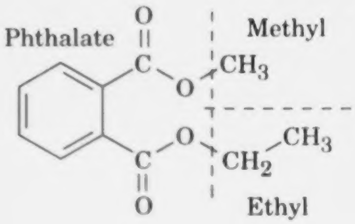
   $$\begin{array}{ccc} H & & H \\ | & & | \\ C & = & C \\ | & & | \\ H & & H \end{array}$$

   Obviously, the expanded formula is clearer, but it takes up space in textbooks that publishers can ill afford to waste. Hence, you will see condensed formulas in print for the most part.

Problem Solving

Routes to Chemical Synthesis

Once students learn a few reactions (oxidation of aldehydes, esterification of carboxylic acids, hydrolysis of esters) it can be fun to write an unfamiliar structure on the blackboard and imagine what you can make with it. If the structure has more than one functional group, it might undergo more than one reaction, enabling you to prepare several interesting compounds. The heart of this synthesis game is recognizing the functional groups and remembering their reaction patterns.



When students have played this synthesis game a few times, it is fun to provide a product and ask them what reactant molecules they need to prepare it. Organic chemists specializing in synthesis do this by taking the desired product and visually breaking it into pieces to see if they can imagine molecules with functional groups that will enable them to prepare the product. (This procedure, called retrosynthesis, is described in detail in most college-level organic chemistry texts.) For example, suppose you asked students to propose compounds to be used in the synthesis of methyl ethyl phthalate.

Although recognizing potential reactants is a trivial problem (see the following note) for an organic chemist, having high school students recognize that methanol, ethanol, and phthalic acid are necessary for the preparation is not at all obvious, even after they have prepared simple esters themselves.

Once students have proposed a reasonable set of starting materials and synthetic steps for the compound, ask them to predict other compounds formed, again based on the functional group properties. In the above example, if one reacted ethanol with phthalic acid first, one would expect to get some diethyl phthalate and unreacted phthalic acid in the mixture as well as the desired product in which one carboxylic acid group had formed an ethyl ester while the other remained unreacted and therefore available to react with methanol. Such pattern recognition is more like solving crossword or picture puzzles than the quantitative problems typically presented to first year chemistry students, but it is critical for all organic chemists. For additional examples, see college-level organic chemistry texts.

[*NOTE: The actual laboratory preparation of a pure sample of methyl ethyl phthalate from phthalic acid, methanol and ethanol is not a trivial problem, even for a synthetic organic chemist. In fact, an organic chemist might proceed from quite different starting materials in order to avoid undesired products like diethyl and dimethyl phthalate. Those who become professional synthetic organic chemists will learn in later coursework some of the problems and exceptions not hinted at in the above exercise. However, the purpose of the above exercise is for students to recognize patterns, and the exercise is valid for that reason.*]

Decision Making

A societal issue related to organic chemistry that students can use for practice in decision-making is petroleum uses. Typically, people think of some sort of fuel when the word "petroleum" is used, but petroleum is also an important chemical feedstock, *i.e.*, it can be used to provide builder molecules for synthesis of drugs, dyes, or plastics important to our lives and lifestyles. Once students have prepared esters, they will be aware that organic chemicals can be used to prepare fragrances and flavors. Chemicals that make excellent fuels—alcohols—are also used for synthesizing fragrances and flavors by reacting with the appropriate carboxylic acid. The hydrocarbon group in alcohols makes these compounds good fuels, but the hydroxy (alcohol) functional group makes them good molecular building blocks.

The same principles apply to compounds derived from petroleum. Functional groups are developed during the cracking process. As petroleum supplies dwindle, society will be forced to decide whether to convert petroleum into CO_2 and H_2O by burning it as fuel, or to convert it into other kinds of molecules—drugs, plastics, dyes, fragrances, flavors, and so forth. An outstanding resource is the *Chemistry in the Community (ChemCom)* curriculum developed under the auspices of the American Chemical Society (1993). A unit in that curriculum is entitled "Petroleum: To Build or to Burn." A shorter resource that makes many of the same points is an article by Seymour (1989)—see *References* section). Also consider the following poem (Kolb & Kolb, 1979) that appears in another excellent article about petroleum chemistry.

Petroleum

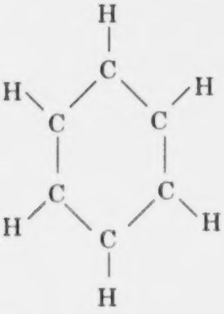
That viscous, tarry liquid nature laid beneath the ground
A hundred million years before people came upon the scene
Can be transformed to marvelous new products we have found—
Like nylon, orlon, polyesters, polyethylene,
Synthetic rubber, plastics, films, adhesives, drugs, and dyes,
And other things, some that we now can only dream about.
Who knows what wondrous products people might some day synthesize
From oil! Except, alas, that our supplies are running out.
The time is near when earth's prodigious flow of oil may stop,
(We've taken so much from the ground with no way to return it);
Meanwhile we strive to find and draw out every precious drop, . . .
And then, incredibly, we take the bulk of it and burn it.

History: On the Human Side

Friedrich August Kekulé (1829-1896) was one of several important German chemists in the 1800's. It was perhaps fortuitous that his fame endures as a great teacher and chemist for he began his university days as a student of architecture and later turned to chemistry after hearing some dynamic lectures by Justus von Liebig.

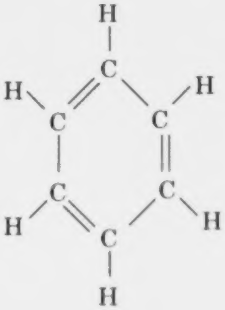
In 1858, Kekulé left his first teaching post at Heidelberg, Germany and went to Ghent, Belgium as professor of chemistry. That same year, both he and Archibald Scott Couper, a Scottish chemist, independently proposed that in all organic compounds, carbon always has four "affinity units," *i.e.*, carbon is tetravalent; it forms four bonds when it joins elements to form compounds. Perhaps it was because the German chemists in the 1800's were receiving wide recognition in organic chemistry while the Scottish chemists were not as well known that fame was heaped on Kekulé rather than Couper for this proposal. Even today, structural formulas where two electron covalent bonds are indicated simply by a line drawn between two atoms are called "Kekulé structures." Before 1858, organic chemists had been working without a mental picture of substances at the molecular level. The Kekulé-Couper theory of molecular structure of carbon compounds shed light on the structural basis of organic chemistry.

Kekulé went on to propose that carbon atoms can bond together to form extended chains of atoms linked together. He is quoted as coming to that conclusion after a dream he had while napping on the top deck of a bus while visiting London. In the dream, he saw carbon atoms uniting in chains and whirling in a giddy dance with smaller atoms (hydrogen, *etc.*) connected to the chains.



Benzene structure proposed by Kekulé

Benzene posed a major problem for chemists in the 1800s because of its unusual properties. The molecular formula was known to be C_6H_6, but no one had been able to suggest a suitable structural formula. In 1865, Kekulé proposed a concept for benzene's structure as presented to him in a dream. While working on a textbook in his study his thoughts wandered, and he dozed before a warming fire, seeing atoms dancing before his eyes—long rows of atoms twining and twisting in snake-like fashion. Suddenly, one of the snakes turned and seized hold of its own tail and the form whirled before his eyes. He awoke and spent the rest of the night working out the cyclic structure of benzene.



Modern structure of Benzene

Some writers have doubted Kekulé's account of the "dream" in developing structural formulas of organic compounds, but none can deny the importance of his work. The development of dyes, drugs such as sulfanilamides and aspirin, high octane gasoline, plastics, fabrics such as Dacron, are all outgrowths of the aromatic chemistry for which Kekulé helped lay the foundations with his structural formula for benzene. Kekulé died in 1896 before the first Nobel prizes were given. However, he is recognized as one of the greatest teachers of chemistry in the nineteenth century. Three of the first five Nobel prizes in chemistry were awarded to his students.

Ira Remsen (1846-1927) was the person most responsible for elevating the reputation of American universities to those of their European counterparts in the late 1800's and early 1900's by incorporating laboratory research into the instruction process.

Born in New York, Remsen was educated in the public schools and went on to receive an M.D. from Columbia University at the age of 21. He was much taken by chemistry, however, after hearing lectures on the subject. Finding no U.S. institution with a reputable graduate chemistry program, Remsen went to Munich to study under Liebig. Eventually, he went to the University of Göttingen to study in the laboratory of Rudolph Fittig, where he received his Ph.D. in 1870.

After two years of postgraduate work in Germany working on the oxidation of substituted benzene rings, Remsen returned to New York to seek an academic appointment. He was named professor of chemistry and physics at Williams College and remained there four years performing his own research, but he was unable to offer coursework in laboratory research for his students. In 1876, at the age of thirty, Remsen was offered a professorship of chemistry at the new Johns Hopkins University in Baltimore. There, he was able to introduce to his teaching the laboratory research component which had influenced him so greatly in Germany. Remsen's approach was so successful that he was able to attract large numbers of students from both the U.S. and Europe, particularly at the graduate level, marking the beginning of the rise in reputation of American institutions in chemistry.

In 1879, an event occurred that was to bring more fame to Remsen's laboratory. Constantine Fahlberg, a Ph.D. from Germany, was invited to work in Remsen's laboratory and continue the study of the oxidation of substituted benzene rings, which Remsen had started in Europe. At Remsen's suggestion, Fahlberg oxidized *ortho*-toluenesulfamide and accidently spilled some of the product on his hand. He noticed it tasted unusually sweet. (Chemists often tasted and inhaled new materials they worked with and were not nearly as cautious as they are today.) Fahlberg saw the importance of this substance as a commercial product and took out a patent on it in 1885. He called it saccharin from the Latin word for sugar, *saccharum*. It had a potency relative to sucrose of about 300 to 1. Remsen apparently felt some grievance about Fahlberg's behavior concerning the patent, and there was some ill will between the two for several years.

The achievements in Ira Remsen's laboratory were many. During this time he founded the American Chemical Journal, the first publication devoted to American chemical research. In 1901, Remsen became the president of Johns Hopkins University while also remaining chairman of the chemistry department. In 1908, he retired from the department but remained president of the institution until 1911. Ira Remsen is renowned as one of the most influential American chemists and educators of the 19th century.

HUMOR: ON THE FUN SIDE

1. **Student response on exam**: Organic chemistry is the chemistry of the organs; inorganic chemistry is the chemistry of the inner organs.
2. **Signs on bumper stickers**:
   a. It takes *alkynes* to make a world.
   b. It is *amino world* without chemists.
   c. Organic chemists do it naturally.
   d. We're *alkene* chemists!
3. **Organic Chem Gems** *(CHEM 13 NEWS, November 1976, p. 1099 and March 1985, p. 5)*
   a. What did the triple-bonded hydrocarbon say to the other hydrocarbons when they threatened to kick him out of gasoline?

      It takes alkynes!

   b. What did the first man to synthesize an organic compound yell when he did it!

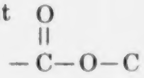
      Urea! I've done it!

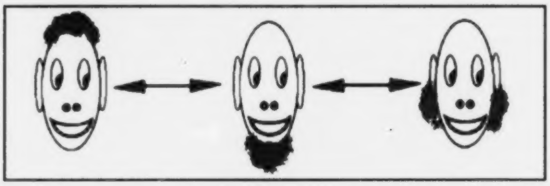
   c. What did the patient say when he was put under?

      Ether that's anesthetic or I'm crazy!

   d. What kind of car does an organic chemist drive?

      A Mercedes Benzene

   e. **Professor**: Well, class, next lecture we will talk about  $-\overset{\overset{\displaystyle O}{\|}}{C}-O-C$ containing compounds.
      Class: You mean. . .?
      Professor: Yes, tomorrow is esterday.
4. **Resonance** *(CHEM 13 NEWS, October 1988, p. 15)*



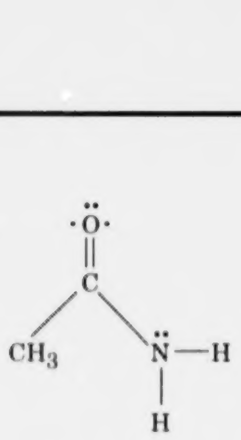
5. **THE ESTER SONG** *(CHEM 13 NEWS, December 1987, p. 4)*

   (To the Oscar Meyer Bologna song)

   My ester has a first name
   that comes from alcohol.
   My ester has a second name
   from acids one and all.
   Ester compounds smell so nice;
   Flavoring our food precise.
   Even though their names are long,
   the world without them would be wrong.

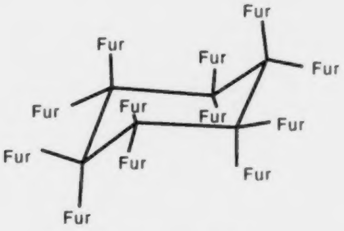
   The trick to naming esters
   is to find the acid part.
   The carbon has two oxygens
   right there in its heart.
   Drop the ic, add A T E.
   It's as easy as can be.
   So we sing our little song;
   our names for esters are not wrong.

6.



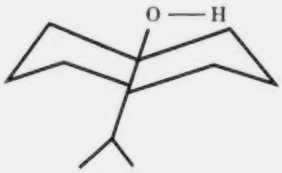
Be very careful with this chemical. It has amide of its own.

CHEM 13 NEWS, October 1985, p. 12



Expensively covered chair

CHEM 13 NEWS, January 1982, p. 8



"Alcoholic" pterodactyl

CHEM 13 NEWS, November 1984, p. 10

7. LAMENT OF THE ORGANIC CHEMISTRY STUDENT

(CHEM 13 NEWS, April 1979, p. 13)

Oh what a pleasure life would be
Without organic chemistry.

No one to take it out of my hide
When I don't know thioacetamide.

Alcohols, acids, phenones, and ketones,
Baffle the student who wails and bemoans;

I study for hours before major tests!
In organic chem. class nobody rests!

While taking this course, one bonds to his book,
That's 1.5 angstroms with a single hook.

The book that we use weighs nearly five pounds.
From twelve hundred pages its mighty word sounds;

What would life be without C, H, and O?
What a disordered place; world full of woe!

Bodies without enzymes, carbohydrates, or fat,
It would be hard to survive without things like that!

To top it all off what would be missed most of all,
That which keeps us all going, Blessed Alcohol!

Straight chains, branch chains, cyclic compounds;
To an ordinary student this surely confounds.

There's a special fraternity of O. C. students,
For together they've suffered through these events.

Though it's only been months, it seems like years,
I've been saturated with O. C. till almost in tears.

The nerve racking tension, the labs that one dreads,
Deep into the soul it seeps and spreads.

You think, eat, and sleep organic chemistry,
Finally synthesizing compounds are all that you see.

Marked for life by this course with which you vie,
By the tremble in your voice, and the glaze in your eye.

It's tough for sure and it won't get any easier;
And as the semester proceeds your stomach gets queasier.

Till the end of the semester you're ready to die,
And hope with a "D" you'll finally slide by.

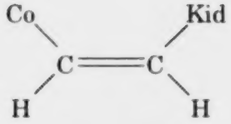
Perhaps the world would be no special place,
Without this science which we all face.

Do you feel that your mind is off in space?
Ah, you're rapidly becoming a basket case!

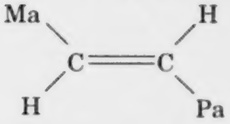
But such is the course of organic chemistry,
Which makes it not for you and me.

8. **Structural puns**. Many puns based on *cis* and *trans* isomerism of aliphatic compounds, *ortho*, *meta*, and *para* isomerism of aromatic compounds, and general nomenclature of organic compounds are possible. Challenge students to make up some new ones and share them with the teaching community by sending them to the *Journal of Chemical Education* or *The Science Teacher*.

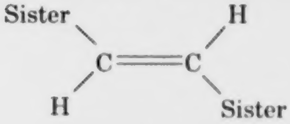
a. ***Cis-trans* isomerism**. (***Cis*** means "on the same side of a double bond." ***Trans*** means "on opposite sides of the double bond."



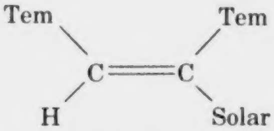
Cisco Kid
(a famous cowboy)



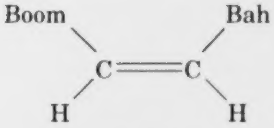
Transparence
(Transparents)


Sister
H
C
C
H
Sister

Electronics
Components
(trans-sisters)

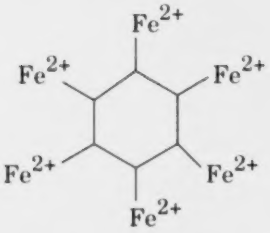

Tem
Tem
C
C
H
Solar

Solar System
(Solar cis-tem)

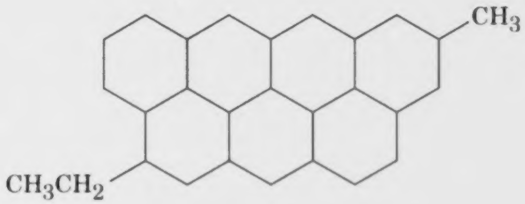

Boom
Bah
C
C
H
H

Football Cheer
(cis-boom bah)

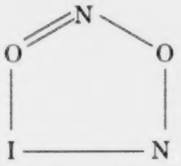
b. **Other**


Fe2+
Fe2+
Fe2+
Fe2+
Fe2+
Fe2+

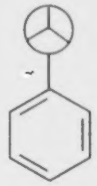
Ferrous wheel
(Older name for iron (II))



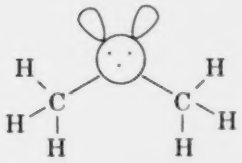
Ethylmethyl chicken wire



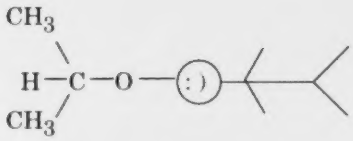
Onion ring



Mercedes
Benzene

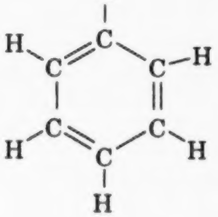


Dimethyl ether bunny



Isopropyl people ether

c. **Aromatic isomerism**. For substituted benzenes, where...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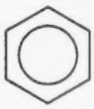
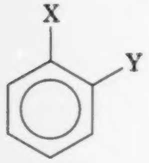
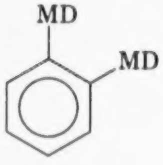
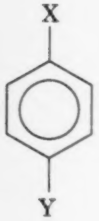




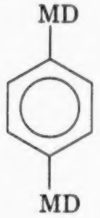
Ortho-



Orthodox
(ortho docs)



Para-



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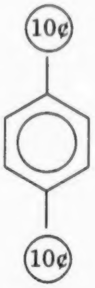
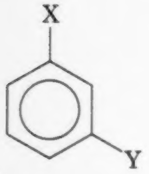
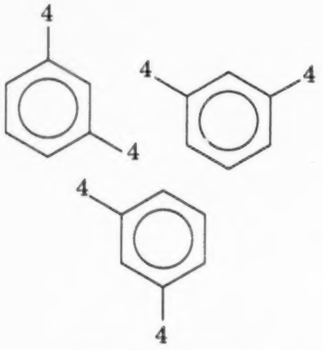
(para docs)



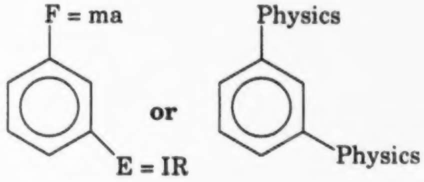
Paradigms
(para dimes)



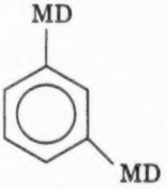
Meta-



Mixed metaphor

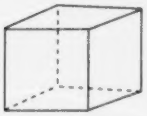

F = ma
E = IR
or
Physics
Physics

Metaphysics



Metaphysicians

d. **SHAPE-DERIVED NAMES**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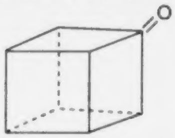
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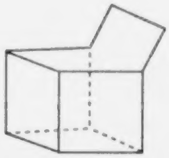
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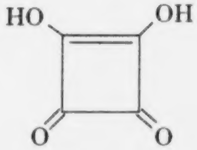
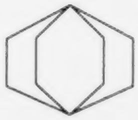
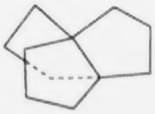
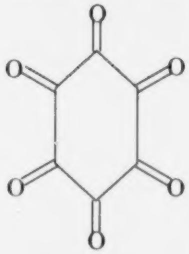
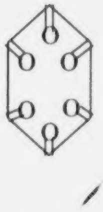

SQUARIC ACID


PADDLANE


PROPELLANE


CYCLONE


ANTICYCLONE

(CHEM 13 NEWS, March-April 1986, p. 8, 12)

9. Word Search (see *Appendix* for master copy)

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L N U G C W N R Q W L N J Y I V X
W Q G T I O V G S N O E R F J M V
G I O L L H V L B B K H W G O D P
H S A N I K G U R F U J T X A E Y
Q O F X H J L A L B L H Y G V E O
Y M Q Y P E C C A B Z G C C U D C
J E M D O O E S V I E P Q K U I G
Y R J I K G E N O N Y W J U H M P
Q K I D D N E S H H Q N X E O A E
D T Y Y Y J U E T H Y L Z L U G A
D H V K H I N O H E S T E R S C A
X Z L O G Q D A O K I Z V P Y G K
F A E J F B U V G B C Y I Z R Q Q
```

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Organic compound that contains a carbon-carbon triple bond.
2. Organic compounds that have pleasant odors and are responsible for odors of fruits and flavorings.
3. Water-loving.
4. Word part meaning "on the same side of a double bond."
5. Another name for chlorofluorocarbons.
6. Predominant type of organic bond in proteins.
7. In a carbonyl compound, this atom is doubly bonded to carbon.
8. Organic compound that contains only carbon and hydrogen.
9. Term that describes structural relationship of ethanol and dimethyl ether.
10. Alkyl group containing two carbon atoms.

Answers: 1. ALKYNE 2. ESTERS 3. HYDROPHILIC 4. CIS 5. FREONS 6. AMIDE 7. OXYGEN 8. HYDROCARBON 9. ISOMER 10. ETHYL

10. Crossword Puzzle (see *Appendix*)
11. See cartoons at end of module.

MEDIA

1. Software published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
   - a. *Alkane Isomers*, by Richard R. Hiatt. Vol. VI B, No. 1, for IBM PS/2 PC-compatible computers.
   - b. *Organic Spectroscopy*, by Richard Hiatt. Vol. II A, No. 2, for the Apple II computer.
   - c. *Proton NMR Spectrum Simulator*, by Kersey Black. Vol. II C, No. 1, for the Apple Macintosh.
2. Software published by Falcon Software, Box 200, Wentworth, NH 03282; (603) 764-5788.
   - a. *SpectraBook for Windows and SpectraDeck for Macintosh – An Annotated Collection of IR, NMR, CMR and Mass Spectra*, by Paul F. Schatz. $95 for individual copies.
   - b. *The Schatz Index for Macintosh* , by Paul F. Schatz. The IR and NMR spectra, physical information, references, synonyms and safety data for 400 organic compounds. $200 for individual copies.
   - c. IR and NMR Simulators of Macintosh and MS-DOS. These simulators are low-cost and easy to use; they are no-maintenance simulations of the real instruments. Data sets for 300 compounds are available separately. $95 for individual copies of NMR simulator, $75 for individual copies of IR simulator, $200 for IR data set, $250 for NMR data set.
   - d. *Organic Reaction Mechanisms*, by J.R. Buell and A.F. Montana. For the Macintosh computer. $75 for individual copies.
   - e. *OGRE–Organic Reaction Drills*, by V.I. Bendall. For IBM computers. $75 for individual copies.
   - f. *Introduction to Organic Chemistry*, by Stanley G. Smith. For IBM PC, IBM PS/2, Apple II and Macintosh computers. 7-9 single disks at $60 per disk.
3. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).
   - a. For the Apple II computer running on ProDOS: AR 702, AR 705.
   - b. For the Apple II computer: AP 701, AP 702, AP 704, AP 705, AP 706, AP 707, AP 717, AP 727.
   - c. For IBM PCs and PC-compatibles: PC 3501, PC 3504.
4. Videodiscs published by*JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
   - a. "Carbon and its Bonds," "Making Aspirin," "Perkin's Experiment," "Chiral Molecules and Polarized Light," "Making Polyethylene," and "Polyethylene," six chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc II (double sided, 60 min.), Special Issue 4.
   - b. *Demonstrations in Organic Chemistry* (double sided, 60 min.). Special Issue 6.
5. World of Chemistry Videocassettes, "Number 21: Carbon" and "Number 22: The Age of Polymers." Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.

EQUIPMENT AND INSTRUMENTATION

Modern organic chemistry is dependent on instrumentation. Organic chemists have extensive laboratory instrumentation at their disposal. These instruments are used to obtain detailed information about the presence of certain functional groups and the overall structure of organic compounds. In fact, chemists routinely use instruments and identify them by their initials—GC for gas chromatography, IR for infrared spectroscopy, and NMR for nuclear magnetic resonance (see *Instrumentation* module).

Links/Connections

WITHIN CHEMISTRY

1. **Halocarbons and environmental chemistry.** Although halogen substituted hydrocarbons pose unique environmental problems, they also provide solutions to a wide range of problems.

   a. **Halocarbon pesticides.** Some of the earliest effective pesticides (*e.g.*, DDT and Chlordane) are halocarbons. Their widespread use has allowed humans to temporarily conquer insect-borne diseases (such as malaria) and to increase agricultural productivity. However, the amount of such halocarbons used and their chemical stability resulted in their being spread throughout the ecosystem. The functional group reactivity that made them toxic for insects also made them toxic for some other species such as birds. It is a misconception that only humans make toxic halocarbons. A number of organisms have been discovered that produce such substances. The problem with halocarbons of human origin is one of scale and concentration, not the "un-naturalness" of the compounds.

   b. **Carcinogenic halocarbons.** Some halogenated hydrocarbons that were industrially useful were later found to be carcinogenic in animals, or to have other negative effects. For example, one kind of halocarbon, polychlorinated biphenyls or PCBs, were used as dielectric fluids in power line transformers until researchers discovered many of them were carcinogenic.

   c. **CFCs and the Ozone layer.** Another kind of halocarbon causing problems recently is the group of compounds known as **freons** or CFCs (chlorofluorocarbons). CFCs are essentially inert at the earth's surface. They have been used as coolants in refrigerators, freezers, and air conditioners. Unfortunately, their lack of reactivity at the earth's surface and their high volatility made it easy for CFCs to float higher up in the atmosphere. Once in the upper atmosphere, they undergo a characteristic reaction of halogen as a functional group—that is, they absorb UV light and break into pieces called free radicals that contain unpaired electrons. The unpaired electrons give the free radical breakdown products of CFC just exactly the right chemical properties to react with ozone (O_3) occurring naturally in the upper third of the atmosphere. Unfortunately, O_3 is needed there to absorb UV radiation that would otherwise reach the earth's surface damaging us and other living systems.

2. **Fuels and Thermodynamics.** Concepts learned by applying thermodynamics to the exothermic reaction of NaOH with HCl can be applied to hydrocarbons. Burning either petroleum or coal supplies most of the world's energy. The enthalpy of reaction for combustion of hydrocarbons can be experimentally determined. For example, see the experiment (*References* section) for determining the heating value of fuels reported by Rettich, Battino, and Karl (1988). Thermodynamic problems can focus on the burning of organic compounds.

3. **Reaction Kinetics**. The majority of synthetic organic reactions occur in solutions at temperatures below 150 °C. In addition, most reactions can be considered complete in just a few minutes. These kinds of chemical reactions lend themselves to kinetic studies. Reactions which consume or liberate acids or bases can be followed by adding a small amount of acid-base indicator to the reaction solution. An excellent example of just such a reaction is the hydrolysis of a methyl ester in a base solution. See *References* section for Rodig, Bell, and Clark (1990). Many organic reactions occur in a stepwise fashion and can be described by a series of equations called a reaction mechanism. One of the best ways of testing a proposed reaction mechanism is to see if it fits an experimentally determined rate equation. Finally, there are numerous examples of catalysis in organic chemistry. Some reactions are catalyzed by species in solution and some are catalyzed by surfaces in contact with solutions.

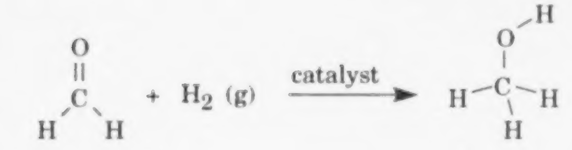
4. **Molecular Geometry**. The geometry of organic compounds is an important consideration in predicting their reactivity (see the discussions of isomerism in *Links and Connections to the Contemporary World*). Most organic compounds have a number of tetrahedral centers, and other centers that are trigonal or linear. Since each atom is bonded to carbon by an electron cloud, counting the number of atoms around a particular carbon atom in a molecule permits prediction of the number of electron clouds around that carbon. VSEPR theory can then be used to predict the geometry around the carbon (see *Molecular Geometry* module). If there are four atoms attached to carbon, there are four clouds, and they form a tetrahedral arrangement. If there are three atoms attached to a carbon (three clouds), they form a planar trigonal arrangement. If there are two atoms (two clouds), there will be a linear arrangement.

5. **Polymer chemistry**. Although progress has been made in recent years to synthesize useful polymers from inorganic materials, most commercially useful polymers are made from organic chemistry builder molecules. In addition to synthetic polymers, there are many naturally occurring polymers based on organic chemicals. DNA and proteins are just two obvious examples. In teaching polymers, emphasize the functional group chemistry that serves to link monomers to form polymers.

   A polymer that students may find interesting is supernylon, which appears ready to assume a dominant role in the marketplace. Stanyl™, developed in the Netherlands, reportedly has a higher melting point and greater toughness and temperature stability than traditional nylon polymers. These properties should make it ideal for the manufacture of molded parts and fibers. For additional information on organic polymers, refer to the *Polymer* module.

6. **Biochemistry**. Biomolecules are organic compounds such as DNA and proteins. Another example is lignins, which provide structural support for plant cells. Large numbers of organic compounds are critical for biochemical functioning. Many of these compounds are mentioned in various contexts in this module.

7. **Biochemical reactions with an aldehyde.** An aldehyde reacts with H_2, in the presence of an appropriate catalyst, to form an alcohol. Hydrogen adds across the carbon-oxygen double bond, placing a hydrogen atom on each of the atoms comprising the carbonyl group.



Formaldehyde

Methanol

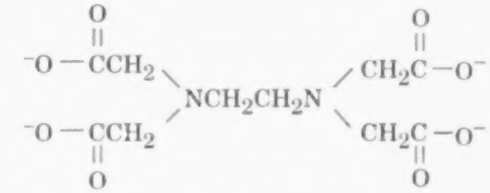
Your body processes alcohol via a mechanism that is the reverse of this process. Whereas adding H_2 to something is called hydrogenation, taking H_2 away is called dehydrogenation. In your body, and in many other species, alcohol is processed by an enzyme called alcohol dehydrogenase. Ingested ethanol reacts with alcohol dehydrogenase in the liver to form acetaldehyde. The aldehydes have characteristic odors. People who have been drinking a lot but quit a few hours earlier have a characteristic odor—not of alcohol but of acetaldehyde, as the body expels it via perspiration and breathing.

Acetaldehyde is reactive because of its aldehyde functional group. It is able to sneak into systems like the ends of nerve cells because of its small size. Acetaldehyde may be the molecule that actually causes drunken behavior rather than alcohol itself.

Aldehydes are very easy to oxidize. The human body oxidizes acetaldehyde to acetic acid. If large quantities of alcohol are ingested, enough may be converted to acetaldehyde which is then oxidized to acetic acid to cause a drop in blood pH. If the pH drops far enough, enzymes necessary for vital functioning begin to precipitate and the individual dies of alcohol poisoning. One of the medical treatments for alcohol poisoning is to add sodium bicarbonate solution to the patient's bloodstream, helping to raise the pH back to normal levels because the bicarbonate anion is the conjugate base of a weak Brønsted-Lowry acid.

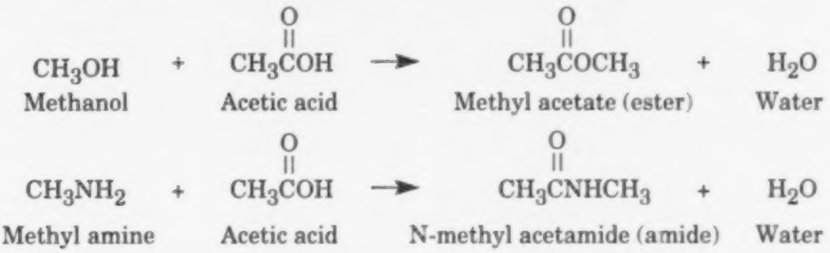
A similar reaction series occurs when methanol (wood alcohol) is ingested because alcohol dehydrogenase can also work on methanol. In this case, formaldehyde is produced by the dehydrogenation of methanol. Formaldehyde is quite reactive and reacts with just about any biomolecule that happens to be in the neighborhood when it is formed. For example, it is known to be a carcinogen, indicating that it can disrupt the cell's genetic machinery. Methanol ingestion, even in fairly small quantities, can lead to blindness, paralysis, and death. Patients who have ingested methanol are treated with ethanol because alcohol dehydrogenase reacts preferentially with ethanol. Maintaining a relatively high concentration of ethanol in the body pushes the equilibrium of the reaction in which enzyme-ethanol complex is formed to the right, effectively preventing oxidation of methanol into formaldehyde. Once analysis of the patient's urine indicates that methanol is no longer being excreted, then the ethanol treatment is removed (see *Chemistry in Medicine* module).

8. **Complex ion formation.** The biological example of hemoglobin as a complex ion in which iron is bound as a Lewis acid to the nitrogen bases in heme was discussed above. Another application of the same principle is the use of EDTA to complex metal ions. EDTA stands for ethylenediamine tetraacetate, and this is its structure.



The four carboxylate groups are free to rotate or flop around, allowing them to adopt a number of different positions in space with respect to the amine nitrogens. The amine nitrogens and the carboxylate groups are Lewis bases that can donate an electron pair to bond with a Lewis acid in complex formation. Consequently, EDTA is able to form complexes with a wide range of metal ions. EDTA is used in analytical chemistry to complex the ion of interest and sometimes to remove a competing ion. It is also used in a variety of other contexts in which forming an organic complex of a metal ion is useful (as in lead poisoning treatment). Have students read food and/or over-the-counter (OTC) drug package labels to see if they can find EDTA as an ingredient.

9. **Formation of amides.** A reaction similar to the formation of esters, described earlier in this module, is the formation of amides. In amide formation, a carboxylic acid reacts with an amine to produce the amide and water. Compare the reactions for ester and amide formation:

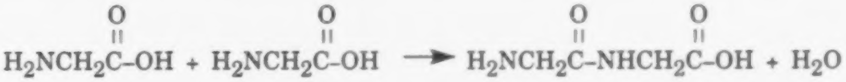


$$\underset{\text{Methanol}}{CH_3OH} + \underset{\text{Acetic acid}}{CH_3\overset{\overset{O}{\|}}{C}OH} \longrightarrow \underset{\text{Methyl acetate (ester)}}{CH_3\overset{\overset{O}{\|}}{C}OCH_3} + \underset{\text{Water}}{H_2O}$$

$$\underset{\text{Methyl amine}}{CH_3NH_2} + \underset{\text{Acetic acid}}{CH_3\overset{\overset{O}{\|}}{C}OH} \longrightarrow \underset{\text{N-methyl acetamide (amide)}}{CH_3\overset{\overset{O}{\|}}{C}NHCH_3} + \underset{\text{Water}}{H_2O}$$

These reactions are very similar. The carboxylic acid functional group reacts in an essentially identical fashion in both reactions. However, the difference in electronegativity between O and N makes the reaction rates and extent of reaction very different. Consequently, the preparation of amides in the laboratory is usually conducted with different starting materials. Biochemically, amide bonds are formed from amines and carboxylic acids with the help of enzyme catalysts.

Amides, like esters, are biologically important and the amide bond is the principal bond in protein formation. Amino acids are organic compounds with two functional groups—an amine group, and a carboxylic acid group. In forming proteins, amino acids are linked *via* an amide bond. An example of a simple amide bond between two amino acids is shown.



$$H_2NCH_2\overset{\overset{O}{\|}}{C}\text{–}OH + H_2NCH_2\overset{\overset{O}{\|}}{C}\text{–}OH \longrightarrow H_2NCH_2\overset{\overset{O}{\|}}{C}\text{–}NHCH_2\overset{\overset{O}{\|}}{C}\text{–}OH + H_2O$$

The newly formed amide still has an amine group on one end and a carboxylic acid group on the other, enabling it to form other amide bonds. The manufacture of polyamides such as nylon is dependent on this principle.

10. **Acids and Bases.** Many applications of organic chemistry to acids and bases exist. A few have been mentioned previously. Some weak acids and bases are organic compounds. Carboxylic acids, *e.g.*, acetic and propionic acids, are typical weak acids, whereas amines, *e.g.*, methylamine and ethylamine, are typical weak bases. In teaching acid-base equilibria, organic as well as inorganic examples can be used.

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

1. **Biology.** The links between organic chemistry and biology and its related disciplines such as medicine are obvious. Specific examples are found throughout this module.
2. **Literature.** Detective fiction utilizes numerous chemical clues. Stories by A. Conan Doyle and Dorothy Sayers are especially good examples. In *A Study in Scarlet* a protagonist is (or at least may have been) slowly poisoned.
3. **Art.** Organic chemistry has a great impact on the art world. Many of the paints and dyes—acrylics, for example—are organic in origin. The instrumental techniques used by organic chemists are often used in determining whether paintings are forgeries. Organic chemistry is also used in color photography.
4. **History.** History is likewise not immune to organic chemistry. The Phoenicians may have based their trading empire on similar technology three millennia ago with their knowledge of the production of dyes and the interaction of dyes with fabrics. It is possible to overemphasize the importance of politicians and military men and to ignore the importance of trade, including new industrial processes, that influence the patterns of history. One of the reasons for Germany's rise to power in the nineteenth century was its industrial prowess in organic chemistry. German companies held so many patents that the American chemical industry found it difficult to grow and prosper. During World War I, the Americans took over many of the German patents and began to manufacture substances like aspirin.

TO THE CONTEMPORARY WORLD

Careers

Organic chemists are employed in large numbers in industry. The pharmaceutical industry employs organic chemists to conduct research to find new and different compounds for medical applications. New areas of research, such as gene splicing, continue to develop. The petrochemical industry has long been a major employer of organic chemists. Many of the top twenty-five chemicals produced in the U.S. are made from petroleum cracking. Fifty percent of industrial chemists in the U.S. work in some area of organic polymer chemistry. Organic chemists are also needed in such fields as environmental protection, the health sciences, including toxicology, the biological sciences, transportation industries and the semiconductor industry. In addition to industrial positions, organic chemists are employed as researchers in government and academic laboratories. Because a basic knowledge of organic chemistry is also required for professionals in such diverse areas as medicine and materials engineering, numerous organic chemists are also employed as teachers. Many organic chemists combine more than one professional description—teacher and researcher, or academic researcher and industrial consultant just to name two possibilities.

Community

Many industries are involved with organic chemistry, from petroleum refining or pharmaceuticals to treatment of wooden poles with creosote, the production of pine oil cleansers, and the treatment of coal or other fuels prior to burning. A potential field trip site includes a visit to the local propane/butane/natural gas distributor, or a swampy area where anaerobic decomposition is producing methane bubbles, or local beauty colleges where knowledgeable representatives of the industry can discuss and demonstrate the various beauty products used. (For example, permanents require the use of an agent that denatures the protein in hair.) You might also wish to contact your county extension agent, who should be knowledgeable in the area of pesticides.

Most areas have a local section of the American Chemical Society. If you don't know where or when yours meets, contact chemistry teachers at the nearest college or university or call the ACS in Washington, DC. By attending a meeting, you can meet chemists in your area who are involved in all sorts of professions. Each local ACS section typically has several organic chemists who will willingly talk with students.

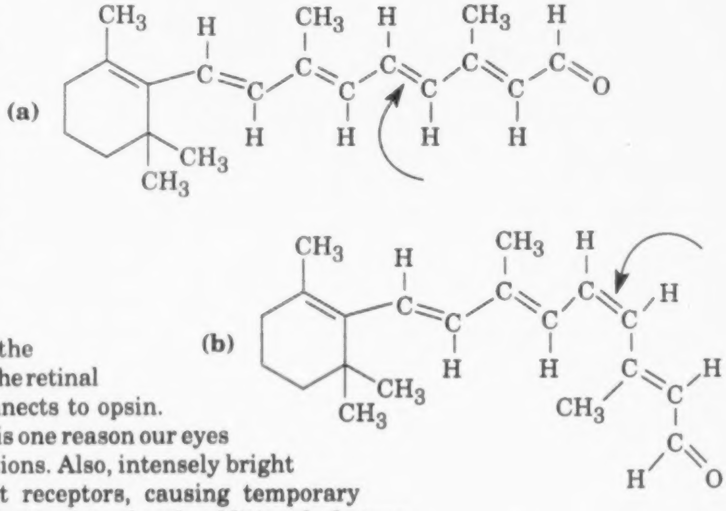
Personal

1. **The Chemistry of Vision involving *Cis* and *Trans* Rhodopsin.**

Vision is actually a very complex chemical phenomenon. When light enters the opening in the front of the eye, it travels through the eye and falls on the rear surface (retina) lined with cone-shaped and rod shaped cells. Each eye contains 7 million cones for detecting color and 120 million rods to detect white light and to provide sharpness of visual images. The molecules responsible for vision are attached to the tops of the rods and cones.

Rhodopsin is one of the molecules responsible for vision. It has two parts: a protein portion called "opsin," and a small aldehyde portion called "retinal." The structure of retinal can occur in the *trans* form where all substituents around the double bonds are *trans*. It can also be noted that vitamin A, a substance known to aid vision—especially night vision, has the same structure as retinal except that the terminal aldehyde functional group is replaced by an –OH group to give an alcohol. Retinal can occur in other isomeric forms, one of which is called the 11-*cis* form because of the *cis* arrangement at carbon 11. It is in this *cis* form that retinal is bound to opsin. When rhodopsin absorbs light, the retinal is isomerized to the all-*trans* form, which separates from the opsin.

The retinal portion of rhodopsin: (a) the *trans* form; (b) the 11-*cis* form.



When the two portions separate, the natural reddish-purple color of rhodopsin is lost. In addition, the cell to which rhodopsin is attached becomes excited. This receptor cell then excites other cells and sends a message to the brain.

After the activation of rhodopsin and the separation of the *trans* form of retinal, the retinal returns to the 11-*cis* form and reconnects to opsin. This process is relatively slow, which is one reason our eyes need time to adapt to low-light conditions. Also, intensely bright light causes saturation of the light receptors, causing temporary blindness since there is no attached retinal to absorb additional photons. [See *References*, *Chemical and Engineering News* (1992, April 6) for the article describing isolation of the visual pigments for the first time.]

2. **Optical Isomerism and Drug Specificity**. Stereoisomers have the same bonds but different spatial arrangements of atoms. One type of stereoisomer is called an optical isomer because the isomers have opposite effects on plane-polarized light. The isomer that rotates light to the right (when viewed down the beam of oncoming light) is said to be dextrorotatory, or *d*. The isomer that rotates the plane of light to the left is called levorotatory or *l*. Optical isomers are mirror images of each other (like left and right hands); they cannot be arranged in space so that they completely overlap.

   Many biomolecules are optical isomers, and their reactions are highly dependent on structure. For example, a drug might have a particular effect because it can bind with a specific isomer in the body. In order to bind correctly, the correct optical isomer of the drug must be administered. Thalidomide is a drug which has two optical isomers. Unfortunately, when the incorrect isomer, or a mixture of both isomers, was administered to pregnant women, it caused terrible birth defects in the children born to these women. (See *Chemistry in Medicine* module for another example—amphetamines.) Many naturally synthesized compounds are a given optical isomer rather than a mixture of the two possible isomers (a 50/50 mixture is called a "racemic mixture"). When a living creature dies, slow racemization of the optically active molecules in the organism begins. Archaeological organic chemists have measured the average rate of this process, thereby developing a method for fixing the date of death of the organism or compounds derived therefrom. This method promises to be a much more sensitive and less limited dating technique than radiocarbon dating.

SCIENCE/ TECHNOLOGY/ SOCIETY

1. **Methane**. Methane, along with carbon dioxide, water vapor and other substances, helps control the earth's temperature. Scientists have noted that the atmosphere's methane content seems to be increasing 2% annually. Although most people don't consider "naturally occurring" methane a problem for society, many of the sources are human-related in origin.

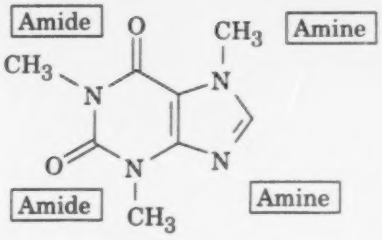
   One source of this increase appears to be termites. When termites consume wood, cellulose is broken down into methane, carbon dioxide and other compounds by microorganisms in termites' digestive systems. It is estimated that termites produce 165 million tons of methane and 55 million tons of carbon dioxide annually. These amounts of methane and carbon dioxide are very significant and are due to environmental changes favoring termite multiplication. One such change is deforestation, which leaves large numbers of felled trees and branches, a perfect environment for termites. The additional methane can contribute to the Greenhouse Effect, possibly increasing global warming.

   Methane gas is also given off by cows as they belch when chewing their cud. In Oregon, this gas has even been collected by some farmers and used as a fuel. Again, that sounds quaint and interesting until one considers the number of cattle raised every year to provide beef, milk, and leather for a growing human population. Bacteria in a cow's stomach is not the only way cattle produce methane. Manure is another methane source. In fact, decomposing manure is such a potent source of methane, it can kill. Experienced farmers in Menominee, MI entered a manure pit and were overcome and later died from methane gas poisoning.

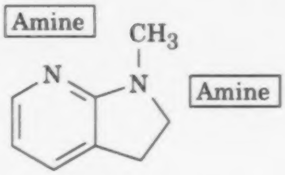
Rice is the major food staple for much of the world, and rice paddies are major sources of methane.

No one knows for sure how much of the methane in the atmosphere results from nonhuman sources and how much results from human activity. It is true, however, that our production of CO_2 and CFCs are not our only contributions to the Greenhouse Effect.

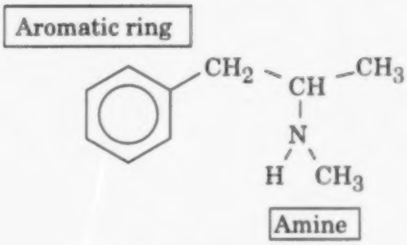
2. **Drugs: Licit and Illicit.** Most drugs, both legal and illegal, are organic compounds, and most of the currently known psychoactive drugs are amines. A few of those currently used by society are caffeine, cocaine, methamphetamine, and nicotine. Important functional groups are indicated in their structures.



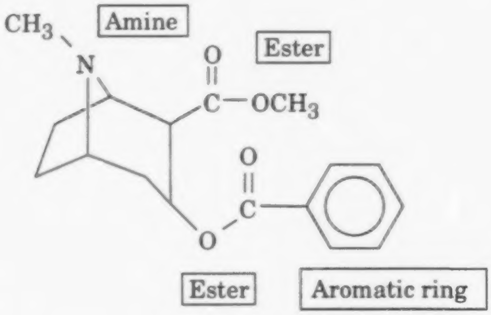
Caffeine



Nicotine

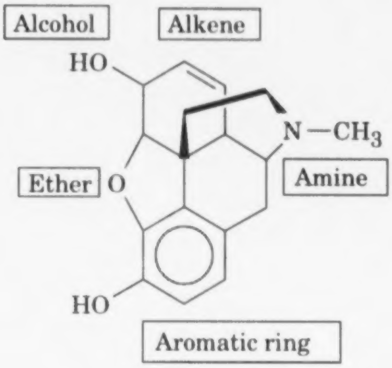


Methamphetamine

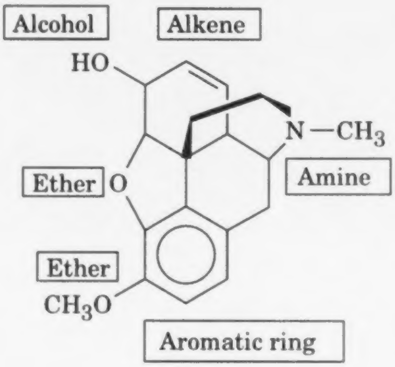


Cocaine

Some scientists have proposed that ethanol, causing most problems of substance abuse in the United States, does not directly cause all symptoms associated with it. Rather, the acetaldehyde formed via alcohol dehydrogenase oxidation of ethanol reacts with one or more naturally occurring brain amines to form products similar to other psychoactive substances in the way they fit receptors on nerve cells. One very well understood group of psychoactive substances that illustrate the way such substances seem to work is the group called the morphine alkaloids. Morphine itself was discovered when chemists began investigating the chemical basis of the pain-killing properties of extracts of the opium poppy. Two naturally occurring cousins of morphine are codeine, which is often used in cough syrups, and thebaine, a pain-killer. Structures for these compounds appear here with important functional groups indicated.

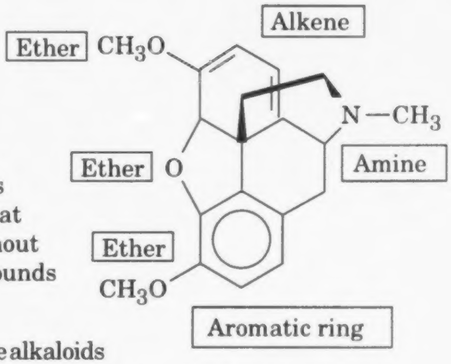

Alcohol
Alkene
HO
N—CH3
Ether
O
Amine
HO
Aromatic ring

Morphine

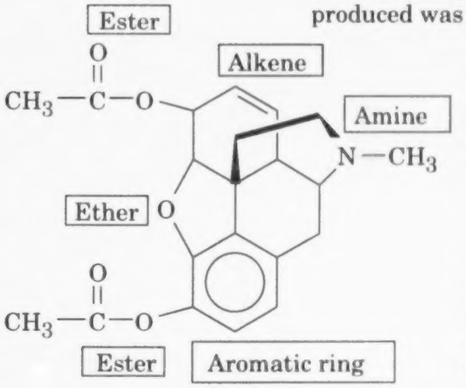

Alcohol
Alkene
HO
N—CH3
Ether
O
Amine
Ether
CH3O
Aromatic ring

Codeine

Although these compounds may be useful in medicine, they have generated social problems because they are addictive. A systematic investigation of derivatives of morphine was carried out in the late nineteenth and early twentieth centuries in an attempt both to understand morphine's mechanism and to produce chemicals that had morphine's pain-killing properties without its addictive properties. One of the compounds produced was heroin.


Alkene
Ether
CH3O
N—CH3
Ether
O
Amine
Ether
CH3O
Aromatic ring

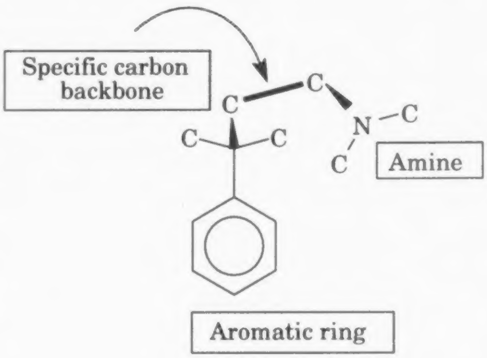
Thebaine


Ester
O
CH3—C—O
Alkene
Amine
N—CH3
Ether
O
O
CH3—C—O
Ester
Aromatic ring

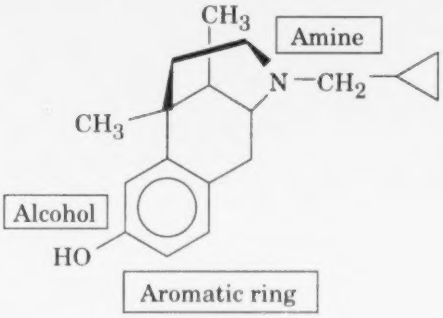
Heroin

Further study of the morphine alkaloids determined that the key structure included an aromatic ring attached to a 3-carbon backbone, the nearest carbon of which was also attached to two carbons and the other end of which was attached to a tertiary amine. (Tertiary amines have three carbons attached to the nitrogen.)

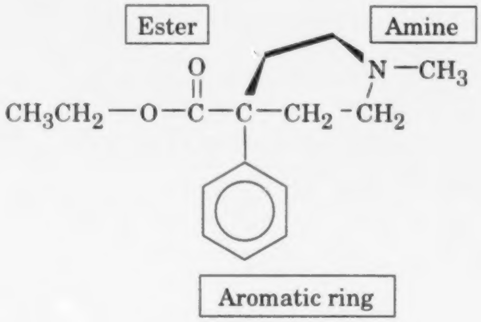
Other research indicated that the mechanism of morphine action was that morphine alkaloids bound to the surface of nerve cells in the brain, in spots called "opiate receptor sites" (named after the original source of morphine alkaloids). The opiate receptor sites normally bind to small proteins such as b-endorphins that occur naturally in the brain. Because morphine alkaloids are a slightly different shape from b-endorphins that usually bind to opiate receptor sites, they alter the shape of the receptor site, thereby altering the way the brain perceives pain. The addictive morphine alkaloids like morphine and heroin also apparently cause long-lasting changes in the brain that result in addiction.


Specific carbon backbone
C
C
N
C
C
C
C
C
Amine
Aromatic ring

From these clues, compounds that have morphine's pain-killing properties but greatly decreased addictive properties were synthesized. Two of these include cyclazocine and meperidine.



Cyclazocine



Meperidine

Examining the structures of all morphine alkaloids, including the nonaddictive synthetic substitutes, shows they fit the general structure allowing binding to the opiate receptor sites. If you have the time to put these structures on the board, you might wish to ask students which feature(s) of morphine, codeine, and heroin might be responsible for the addictive properties. Such a question would comprise a good exercise in pattern recognition.

3. **Pesticides.** A serious problem in providing food for the world's populations is pests: weeds, fungi, insects, worms, rodents, *etc*. Two modes of attack have been made on this problem: breeding hardier plant species and developing chemical pesticides.

A pesticide that kills weeds is an herbicide. The first chemical herbicides were inorganic compounds, but they often killed the crop they were intended to protect. In the early 1940's, a major breakthrough in weed control occurred with the discovery of 2,4-dichlorophenoxyacetic acid or 2,4-D. This substance selectively kills broad leaf weeds. 2,4-D has been widely used to control weeds in wheat, rice, corn and sugar cane, as well as in lawns. A close relative is 2,4,5-trichlorophenoxyacetic acid or 2,4,5-T, first marketed in 1948 and used in controlling brush along roads and fences. 2,4,5-T has been the focus of controversy because of its use in Vietnam to defoliate trees to destroy the cover of enemy soldiers. The substance used was actually a 50:50 mixture of the butyl esters of 2,4-D and 2,4,5-T and was called Agent Orange because it was stored in orange drums. Agent Orange has been implicated in causing a wide range of symptoms in humans. Many American soldiers were exposed to Agent Orange during the spraying, and some of them later joined in suing the U.S. government for damages suffered as a result. Many scientists feel that any damage to humans as a result of the spraying was probably due to an impurity, dioxin, in the mixture. Even so, restrictions have been placed on the use of 2,4,5-T because of the possible toxic effect on humans. Whether Agent Orange caused damage to U.S. soldiers and others is still under dispute. Other herbicides, such as atrazine, alachlor and trifluralin are now being used for similar applications. The herbicides act by interfering with mechanisms important to plant survival. Atrazine disrupts photosynthesis, alachlor inhibits the synthesis of proteins and trifluralin interferes with cell division.

Compounds that kill insects are known as insecticides. Approximately one-third of agricultural produce grown worldwide is consumed or destroyed by insects. The first insecticides were inorganic compounds but most modern products are organic compounds. The first highly successful insecticide was DDT. However, numerous environmental problems were associated with DDT and other chlorinated hydrocarbons. They are toxic and remain a long time in the ecosystem without being chemically degraded to less toxic forms. Chlorinated hydrocarbons have gradually been replaced by other classes of compounds such as organic phosphates, which are much more readily biodegraded. Most of these compounds are closely related to nerve gases developed for human warfare and are quite toxic. Parathion is very toxic to humans, but malathion is less toxic and is widely used because it biodegrades in a few weeks. Another class of insecticides is derived from carbamic acid. The carbamates, of which the compound Sevin is an example, are generally safer to handle but more expensive to produce.

Extensions

1. Compounds containing the carbonyl functional group, C=O (the remaining bonds are to hydrogen or alkyl groups), are called aldehydes (RCHO) or ketones (RCOR'). The chemistry of these compounds is primarily due to the chemistry of the carbonyl functional group. Most aldehydes and ketones give a yellow-to-red precipitate when mixed with 2,4-dinitrophenylhydrazine. However, only aldehydes will reduce Ag^+. The Tollens' test for aldehydes (silver mirror) is based on the reduction of Ag^+ by an aldehyde. (Directions for the Tollens' test can be found in most organic chemistry laboratory manuals; see *References*.) Methyl ketones (ketones where one alkyl is a methyl group) will react with iodine in a basic solution to give a yellow iodoform precipitate. By this difference in behavior you can classify and differentiate aldehydes and ketones—see Pavia, Lampman, Kriz, and Engel (1990) in *References*. Students can devise a set of chemical tests to identify various aldehydes and ketones. In recommending this extension to students, pay special attention to safety features.

2. An excellent independent study project would be to investigate the oxidation of alcohols by $KMnO_4(s)$. Apparently the reactions on the solid surface are different from permanganate oxidations in solution. Parameters not fully investigated include potential for transition metal catalysis (Cu^{2+} has been studied, but not in depth), inhibition by carboxylic acids, and the role of trace quantities of water. Students could isolate the ketones by precipitating as dinitrophenylhydrazone derivatives—see Fieser and Williamson (1979, p. 354) in *References*—identify them by melting point, and/or determine the amount produced per unit time by determining the number of moles of dinitrophenylhydrazone produced. If students had access to a gas chromatograph (most analytical and university laboratories have gas chromatographs), following the course of a reaction would be even easier. Menger (1982) has written an excellent review article for interested students.

3. When alcohols are mixed with solutions of cerium salts, colored complexes (typically orange or red) are often formed. Can the rate of the reaction be studied with a spectrophotometer like a Spectronic 20, a Coleman Junior, a Sargent-Welch Chem Anal, *etc.*? Can students devise a method to follow the reaction by its color change even if they don't have access to a spectrophotometer? Are there any patterns to the colors formed or the reaction rates that can be explained because the alcohols are primary, secondary, or tertiary alcohols, or by the number of carbons in the backbone? Do secondary alcohols that are linear behave differently from secondary alcohols that are cyclic? Can the complex (probably a solid) be isolated? A recipe for the cerium solution is found in Fieser and Williamson (1979), see *References*. They recommend 90 g cerium(IV) ammonium nitrate, $Ce(NH_4)_2(NO_3)_6$ in 225 mL 2 M HNO_3. The idea for this project came from Dr. Archie Moore, Southeastern Louisiana State University, Box 385, Hammond, LA 70402.

4. Do microwaves affect hydrolysis reactions? The production of soap (see *Demonstration 2*) is a hydrolysis reaction. In the demonstration, the reactants were heated on a hot plate. Recent evidence indicates that heating in a microwave oven may cause effects that are not due just to the heat—see Jahngen, *et al.* (1988) in *References*. They reported that microwave heating caused a compound to hydrolyze as much as ten times faster than equivalent hot plate heating. The authors suggest that this "nonthermal" heating may account for observed damage by microwaves to biological systems. Students could replicate Jahngen's study by seeing whether microwave oven treatment produced soap faster than hot plate treatment. They would need to be careful in setting up experimental controls to ensure that thermal heating was equivalent. Many replicates of experiments would be necessary for appropriate statistical analysis of the data.

5. Students who are interested in soaps can prepare them with potassium hydroxide, KOH, as well as with sodium hydroxide, NaOH, using the directions in *Demonstration 2*. In order to compare properties all conditions would have to be carefully duplicated.

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Module developed by William Becker, Beatrice Epperson, and William Lamb, the Oregon team.

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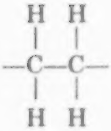
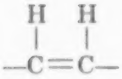
Appendix

- **Transparency Masters**
    1. Common Functional Groups in Organic Compounds
    2. Esters
    3. Organic Chemistry Crossword Puzzle
    4. Word Search
- **Humor**

Common Functional Groups in Organic Compounds


$-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-$

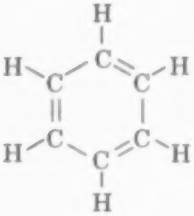
Alkane

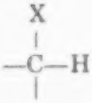

$-\overset{H}{C}=\overset{H}{C}-$

Alkene

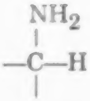
$-C\equiv C-$

Alkyne

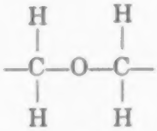

Benzene


$-\underset{|}{\overset{X}{C}}-H$

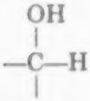
X = F, Cl, Br, I
Alkyl halide


$-\underset{|}{\overset{NH_2}{C}}-H$

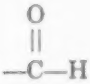
Primary Amine


$-\underset{H}{\overset{H}{C}}-O-\underset{H}{\overset{H}{C}}-$

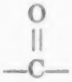
Ether


$-\underset{|}{\overset{OH}{C}}-H$

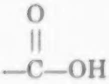
Alcohol


$-\overset{O}{\overset{\|}{C}}-H$

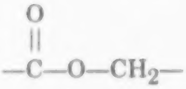
Aldehyde


$-\overset{O}{\overset{\|}{C}}-$

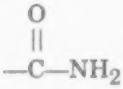
Ketone


$-\overset{O}{\overset{\|}{C}}-OH$

Carboxylic acid

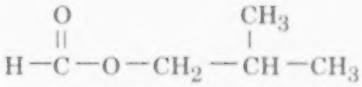

$-\overset{O}{\overset{\|}{C}}-O-CH_2-$

Ester


$-\overset{O}{\overset{\|}{C}}-NH_2$

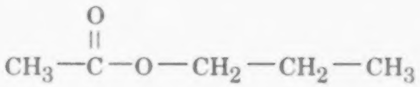
Amide

Esters



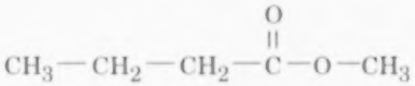
$$H-\overset{\overset{O}{\|}}{C}-O-CH_2-\overset{\overset{CH_3}{|}}{CH}-CH_3$$

Raspberry



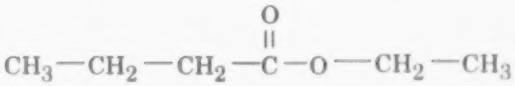
$$CH_3-\overset{\overset{O}{\|}}{C}-O-CH_2-CH_2-CH_3$$

Pear



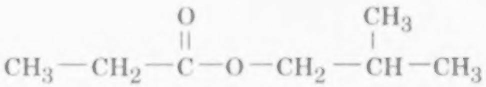
$$CH_3-CH_2-CH_2-\overset{\overset{O}{\|}}{C}-O-CH_3$$

Apple



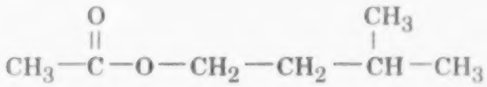
$$CH_3-CH_2-CH_2-\overset{\overset{O}{\|}}{C}-O-CH_2-CH_3$$

Pineapple



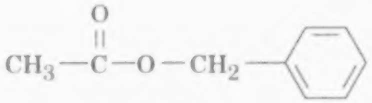
$$CH_3-CH_2-\overset{\overset{O}{\|}}{C}-O-CH_2-\overset{\overset{CH_3}{|}}{CH}-CH_3$$

Rum

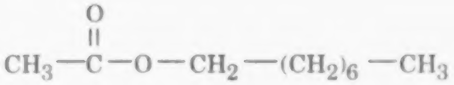


$$CH_3-\overset{\overset{O}{\|}}{C}-O-CH_2-CH_2-\overset{\overset{CH_3}{|}}{CH}-CH_3$$

Banana

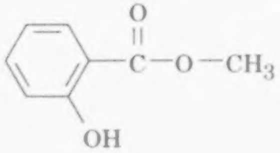


Peach

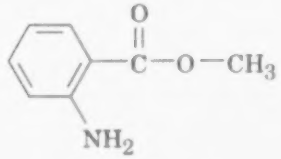


$$CH_3-\overset{\overset{O}{\|}}{C}-O-CH_2-(CH_2)_6-CH_3$$

Orange

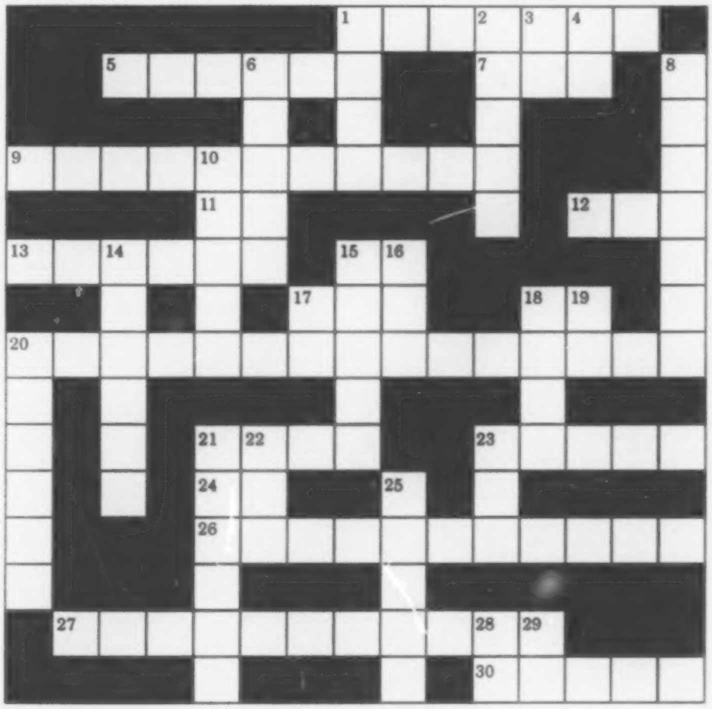


Wintergreen



Grape

Organic Chemistry Crossword Puzzle



Across

1. Saturated hydrocarbons.
5. One of several species with the same formula but with different structure and properties.
7. Configuration applied to geometric isomers denoting an arrangement in which the substituent groups are on the same side of the double bond.
9. Type of organic compound with one or more double or triple bonds.
11. Common verb containing the symbols for Elements 53 and 16.
12. Type of organic compound associated with obesity.
13. Alkyl ______, type of compound in which a hydrocarbon group is bonded to a halogen.
15. Public transportation in Chicago.
17. Country where the organic compound, nylon, was invented (Abbrev.).
18. Suffix for an alcohol.
20. Type of bonding where electrons are shared by atoms and typical of organic compounds. (2 words)
21. Pitcher containing the symbols for Elements 74 and 68.
23. Cyclic, unsaturated hydrocarbon such as benzene.
24. Element 73 (Symbol).
26. Compound composed of carbon and hydrogen.
27. Adjective for compound not soluble in water.
30. Configuration applied to geometric isomers denoting an arrangement in which the substituent groups are on opposite sides of the double bond.

Down

1. Product of length and width.
2. Organic compounds with the general formula RCOOH.
3. Element 28 (Symbol).
4. Element 99 (Symbol).
6. Device used to control certain types of computers.
8. ________ point, temperature at which a pure organic compound is suddenly transformed from a solid to a liquid.
10. Pertaining to the periodic ebb and flow of ocean waters.
14. Water ________, description of a hydrophilic compound.
15. Organic compound produced by the reaction between a carboxylic acid and an alcohol.
16. Where the synthesis in 15-Down takes place, for short
17. Prefix meaning "not."
18. Physical property characteristic of 15-Down.
19. Element 3 (Symbol).
20. Element that all organic compounds must contain.
21. Organic compounds containing a bridging oxygen between hydrocarbon groups and formerly used as an anesthetic.
22. Path or road.
23. Pie ________ mode.
25. Element contained in boric acid.
28. Common pronoun.
29. Element 24 (Symbol).

Across		Down			
1	Alkanes	1	Area	17	Un
5	Isomer	2	Acids	18	Odor
7	Cis	3	Ni	19	Li
9	Unsaturated	4	Es	20	Carbon
11	Is	6	Mouse	21	Ethers
12	Fat	8	Melting	22	Way
13	Halide	10	Tidal	23	A La
15	El	14	Loving	25	Boron
17	USA	15	Ester	28	It
18	Ol	16	Lab	29	Cr
20	Covalent bonding				
21	Ewer				
23	Arene				
24	Ta				
26	Hydrocarbon				
27	Hydrophobic				
30	Trans				

Word Search

L	N	U	G	C	W	N	R	Q	W	L	N	J	Y	I	V	X
W	Q	G	T	I	O	V	G	S	N	O	E	R	F	J	M	V
G	I	O	L	L	H	V	L	B	B	K	H	W	G	O	D	P
H	S	A	N	I	K	G	U	R	F	U	J	T	X	A	E	Y
Q	O	F	X	H	J	L	A	L	B	L	H	Y	G	V	E	O
Y	M	Q	Y	P	E	C	C	A	B	Z	G	C	C	U	D	C
J	E	M	D	O	O	E	S	V	I	E	P	Q	K	U	I	G
Y	R	J	I	R	G	E	N	O	N	Y	W	J	U	H	M	P
Q	K	I	D	D	N	E	S	H	H	Q	N	X	E	O	A	E
D	T	Y	Y	Y	J	U	E	T	H	Y	L	Z	L	U	G	A
D	H	V	K	H	I	N	O	H	E	S	T	E	R	S	C	A
X	Z	L	O	G	Q	D	A	O	K	I	Z	V	P	Y	G	K
F	A	E	J	F	B	U	V	G	B	C	Y	I	Z	R	Q	Q

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Organic compound that contains a carbon-carbon triple bond.
2. Organic compounds that have pleasant odors and are responsible for odors of fruits and flavorings.
3. Water-loving.
4. Word part meaning "on the same side of a double bond."
5. Another name for chlorofluorocarbons.
6. Predominant type of organic bond in proteins.
7. In a carbonyl compound, this atom is doubly bonded to carbon.
8. Organic compound that contains only carbon and hydrogen.
9. Term that describes structural relationship of ethanol and dimethyl ether.
10. Alkyl group containing two carbon atoms.

"Relax Professor relax, you'll be ok. You gave me an A in your organic course."

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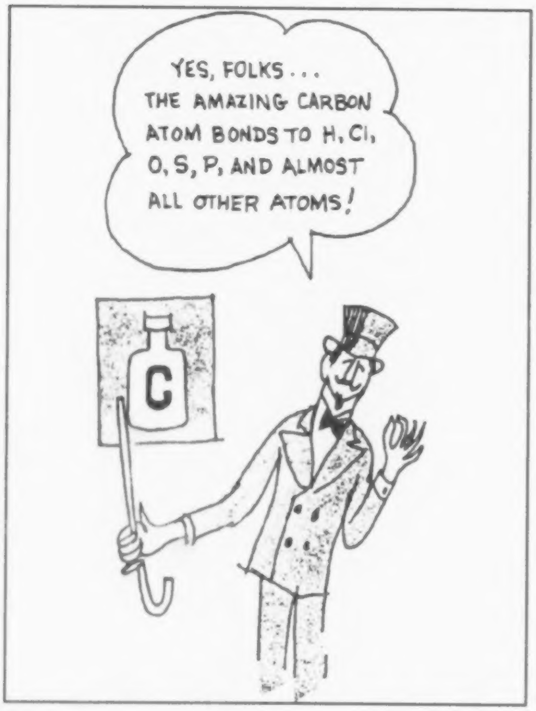
YES, FOLKS... THE AMAZING CARBON ATOM BONDS TO H, Cl, O, S, P, AND ALMOST ALL OTHER ATOMS!
C

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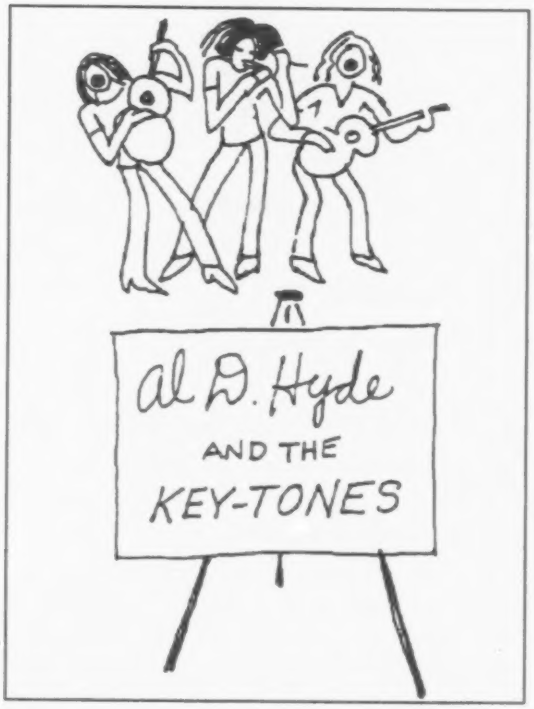
Al D. Hyde
AND THE
KEY-TONES

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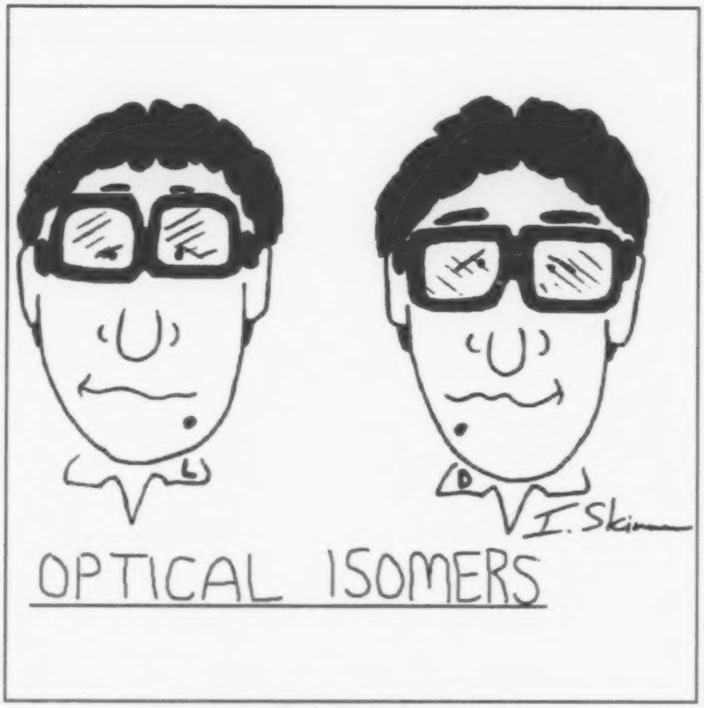
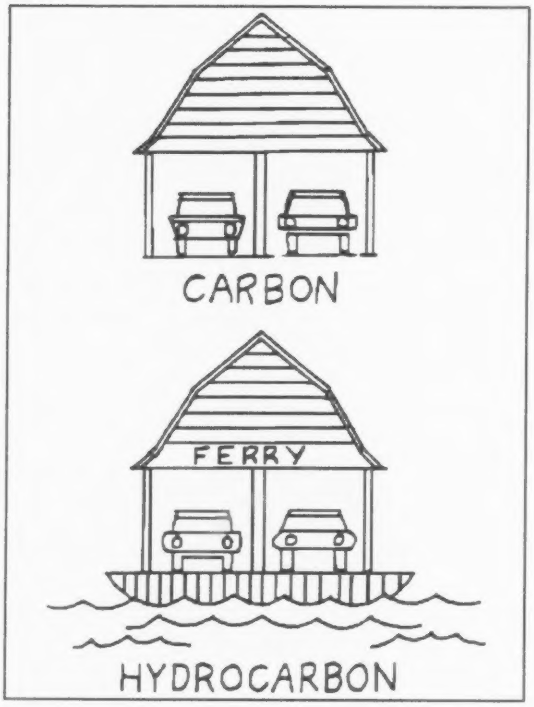

L
D
OPTICAL ISOMERS

CHEM 13 NEWS, September & October 1987, p. 25. Reprinted with permission.


CARBON
FERRY
HYDROCARBON

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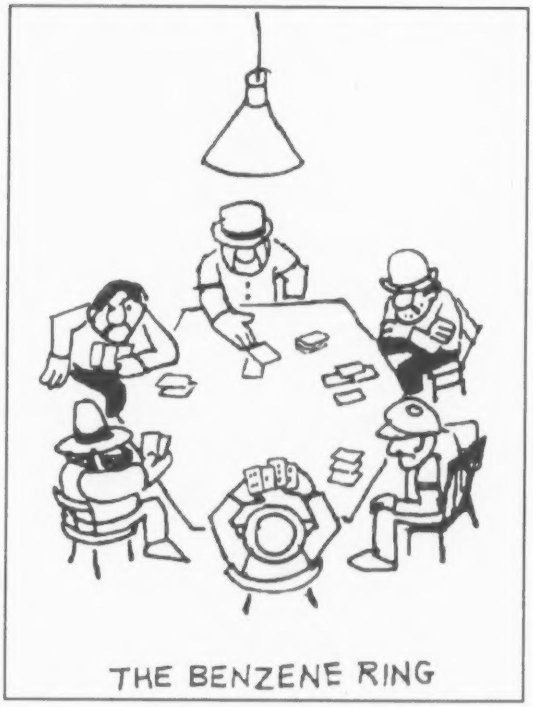
THE BENZENE RING

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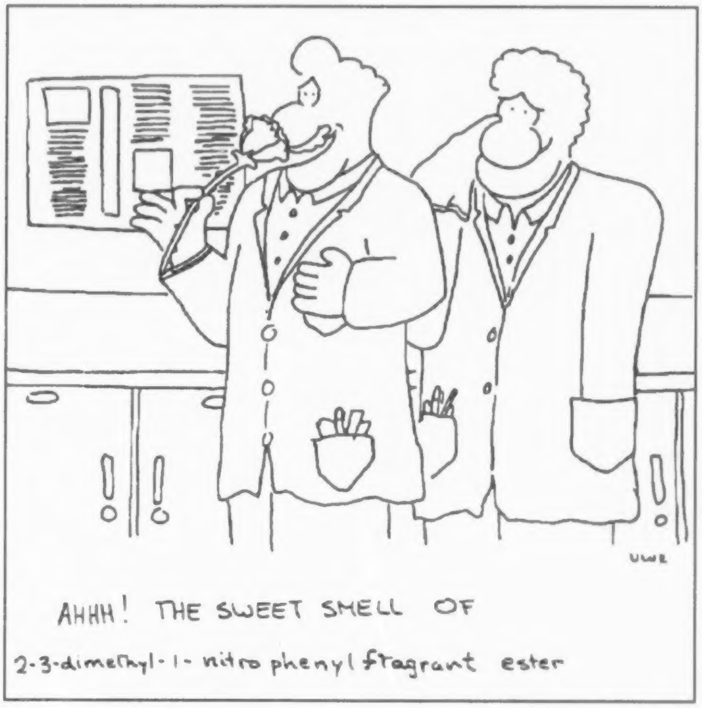
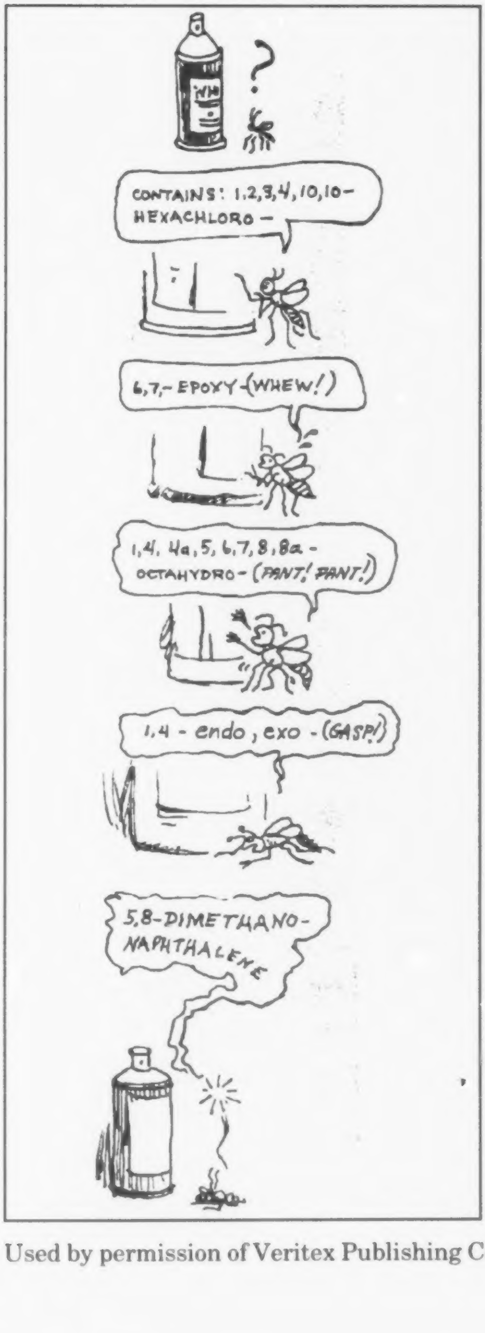

UWE
AHHH! THE SWEET SMELL OF
2-3-dimethyl-1-nitro phenyl fragrant ester

CHEM 13 NEWS, February 1979, p. 11. Reprinted with permission.


?
CONTAINS: 1,2,3,4,10,10-HEXACHLORO-
6,7,-EPOXY-(WHEW!)
1,4, 4a,5,6,7,8,8a-OCTAHYDRO-(PANT! PANT!)
1,4 - endo, exo - (GASP!)
5,8-DIMETHANO-NAPHTHALENE

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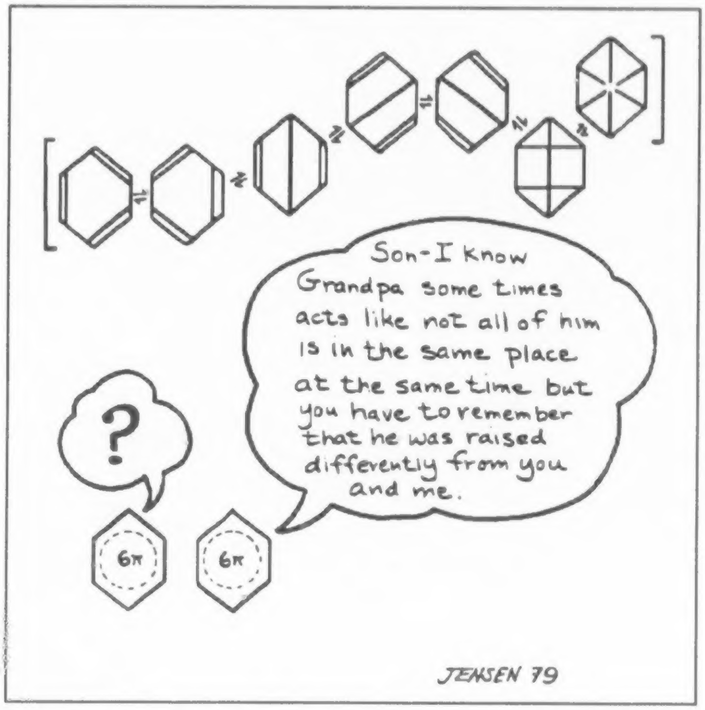
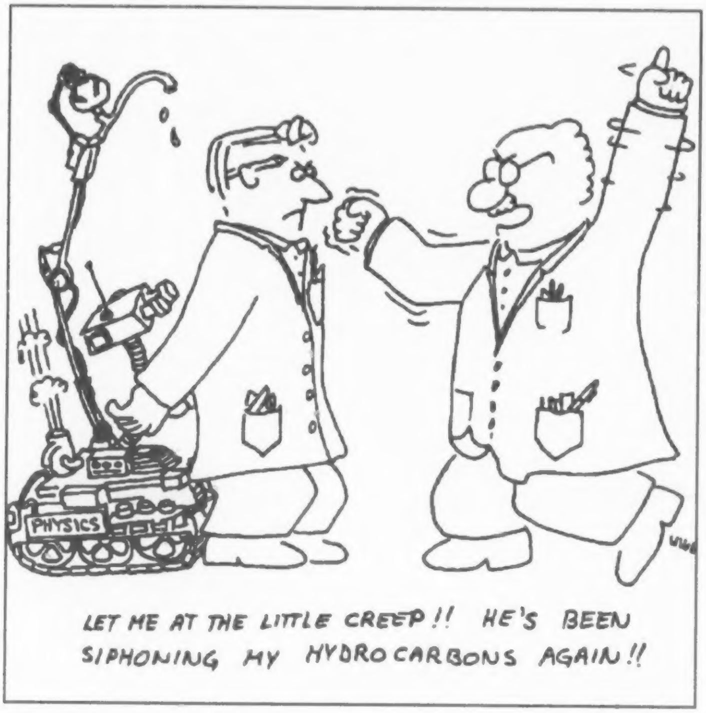
LIVING ON THE EDGE
IN A
BUCKYBALL WORLD

Danylo Shepelavy, U. Penn, 1993. Reprinted with permission.


Son-I know Grandpa some times acts like not all of him is in the same place at the same time but you have to remember that he was raised differently from you and me.
?
6π
6π
JENSEN 79

CHEM 13 NEWS, October 1979, p. 7. Reprinted with permission.


PHYSICS
LET ME AT THE LITTLE CREEP!! HE'S BEEN SIPHONING MY HYDROCARBONS AGAIN!!

CHEM 13 NEWS, March 1980, p. 1. Reprinted with permission.

A SourceBook Module

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OXIDATION-REDUCTION

ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

Oxidation-reduction reactions, or redox reactions, are common. Redox reactions are imagined to involve a *transfer of electrons*. *Oxidation* is always accompanied by *reduction*. Redox reactions are responsible for respiration supporting life, combustion of gasoline, browning of a freshly cut apple, corrosion of metals, operation of an automobile battery, photosynthesis, and many other reactions.

A simple, readily demonstrated example of oxidation-reduction is the reaction of an iron nail with copper(II) ions in a water solution of copper(II) sulfate:

$$\mathrm{Fe(s)} + \mathrm{CuSO_4(aq)} \rightarrow \mathrm{FeSO_4(aq)} + \mathrm{Cu(s)}$$

The process can be *imagined* to consist of two steps. The iron atoms lose electrons and become oxidized:

$$\mathrm{Fe(s)} \rightarrow \mathrm{Fe^{2+}(aq)} + 2e^-$$

The copper(II) ions gain electrons and become reduced:

$$\mathrm{Cu^{2+}(aq)} + 2e^- \rightarrow \mathrm{Cu(s)}$$

The mental picture is one of electron transfer. Oxidation, a loss of electrons, is always counterbalanced by an equal extent of reduction, a gain of electrons.

Another example comes from the formation of magnesium oxide, MgO, from the elements magnesium and oxygen.

$$\mathrm{Mg(s)} + {}^1\!/_2\mathrm{O_2(g)} \rightarrow \mathrm{MgO(s)}$$

Think about the model used to explain ionic bond formation. Magnesium loses 2 e^- to oxygen, thus forming Mg^{2+} ion, while oxygen gains 2 e^- to become an oxide ion, O^{2-}. We always picture a transfer of electrons in oxidation-reduction reactions. In this case, Mg loses two electrons while oxygen gains two electrons. Losing electrons is oxidation; gaining electrons is reduction.

When an atom either in the elemental state or as an ion loses one or more electrons, it is said to be *oxidized*. Conversely, when the atom gains one or more electrons, it is said to be *reduced*.

Also, when an atom is covalently bonded to other atoms in a molecule or ion, we sometimes describe the resulting partial charge build-up on that particular atom in terms of electron loss (oxidation) or electron gain (reduction). Thus, in a reaction such as

$$\mathrm{CO} + {}^1\!/_2\mathrm{O_2} \rightarrow \mathrm{CO_2}$$

the carbon atom is said to be oxidized because two more of its electrons are shared (unevenly) with the more electronegative oxygen.

The term oxidation once implied chemical reaction with oxygen. Today the term is broadly used to suggest loss of electrons during reaction.

In many cases, it is impossible to say that a particular atom has gained or lost electrons. For this reason, a way to characterize redox reactions is based upon *oxidation numbers*. An oxidation number is a positive or negative number assigned to an atom according to an arbitrary set of rules. Rules to determine oxidation states will be developed in this module.

In the reactions used as examples, the oxidation number of each iron atom is changed from 0 to +2, while the oxidation number of each copper atom is changed from +2 to 0. The oxidation number of each magnesium atom is changed from 0 to +2, while the oxidation number of each oxygen atom is changed from 0 to –2. Using rules given later, the oxidation number of each carbon atom changes from +2 to +4 as CO becomes CO_2.

The definition of reduction is often confusing to students. Originally, reduction referred to obtaining metals (elements) from their ores (compounds). Reduction brings about a decrease in oxidation number.

In a redox reaction, the sum of the increases in oxidation numbers for the oxidized atoms is balanced exactly by the sum of decreases in oxidation numbers for the reduced atoms. This fact can be utilized to balance coefficients in any redox equation. Substances in redox reactions that undergo oxidation are called *reducing agents* while substances that become reduced are called *oxidizing agents*. An oxidizing agent is an *agent* that brings about oxidation and, in so doing, becomes reduced.

PLACE IN THE CURRICULUM

This material may follow coverage of periodicity, atomic structure, and conservation of mass and charge. It usually follows stoichiometry, and the first activity requires use of stoichiometry concepts. It should precede any discussion of electrochemistry.

CENTRAL CONCEPTS

1. Many chemical reactions involve the apparent transfer of electrons between reacting chemical species.
2. Oxidation occurs when a reacting species loses electrons; reduction occurs when a reacting species gains electrons.
3. Oxidation is always accompanied by reduction.
4. The energy released from redox reactions can be much larger than that released from acid base reactions; in that case reactant bonds are relatively weak while product bonds are extremely strong.
5. Some oxidation-reduction reactions occur slowly while others occur very rapidly.
6. Oxidation numbers can be assigned to atoms according to a set of rules.
7. A redox reaction always involves a change in the oxidation number of two or more atoms during a reaction.
8. Oxidation numbers can be used as a "bookkeeping" device in balancing redox equations, predicting further reactivity, and in identifying species that exhibit multiple oxidation states. They are assigned on the basis of associating shared electrons with the more electronegative element.
9. Oxidizing agents are reduced; reducing agents are oxidized.
10. Oxidation-reduction reactions permeate daily life. Examples of common redox reactions range from respiration and combustion to corrosion and to the use of anti-oxidants in foods.

RELATED CONCEPTS

1. Conservation of mass and charge
2. Electronegativity
3. Periodic Table and periodicity of properties
4. Equation balancing by inspection
5. Stoichiometry
6. Bond formation theory

RELATED SKILLS

1. Predict common oxidation numbers using the Periodic Table.
2. Recall charges on common polyatomic ions.
3. Use an algebraic number line.
4. Predict bond formation using electron configurations and Lewis electron dot structures.

PERFORMANCE OBJECTIVES

After completing their study of oxidation-reduction, students should be able to:

1. define using their own words, and with examples, the terms: oxidation, reduction, redox, oxidizing agent, reducing agent, oxidized, reduced.
2. compare the quantity of energy released from the most vigorous redox reactions to that released from the most vigorous acid-base reactions.
3. determine the oxidation numbers of atoms in both reactants and products in simple reactions, identify the oxidizing and reducing agents, and identify the substances oxidized and reduced.
4. use the concept of oxidation number together with the concept of a number line to illustrate how oxidation and reduction always occur to the same extent.
5. cite examples in which redox reactions are involved in biochemistry, production of common goods, and environmental changes.
6. balance simple oxidation-reduction equations by at least one method (oxidation number or half-reaction method.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Oxidation States of Manganese

(Modified from Experiment C15 of Doing Chemistry*)*

Introduction

Atoms in compounds may be found in various oxidation states. Atoms of some elements, like manganese, have many possible oxidation states. Oxidation-reduction reactions lead to changes in oxidation states.

Purpose

The oxidation state of manganese in potassium permanganate is changed through a series of redox reactions. The redox reactions are controlled by adjusting the acidity of the reaction medium.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Dilute sulfuric acid and dilute sodium hydroxide solutions are corrosive. Solid potassium permanganate is a strong oxidant. Potassium permanganate and sodium bisulfite are toxic.
3. Use caution when handling the dilute sulfuric acid solution.
4. Do not permit the potassium permanganate to come in contact with any reducing agents. Save any spilled material for disposal according to the procedure given.
5. Dispose of all materials as your teacher directs.

Procedure

Use pulled Beral pipets

1. Put 10 drops $KMnO_4$ solution in four wells of a 24-well plate, A-1 through A-4.
2. Add four drops NaOH solution to well A-2.
3. Add three drops 3M H_2SO_4 solution to well A-4.
4. Place the plate on a sheet of white paper or other white background.
5. Add three drops $NaHSO_3$ solution to well A-2 while stirring with a clean toothpick until there is evidence of a reaction.
6. Add three drops $NaHSO_3$ solution to well A-3 while stirring with a toothpick. Note evidence for a reaction.
7. Slowly add $NaHSO_3$ solution to well A-4, while stirring with a toothpick, until there is evidence of a reaction.
8. MnO_2 is an insoluble brown solid. Mn^{2+} is faintly pink in solution. MnO_4^{2-} is dark green in solution. Based upon these colors, identify the manganese reaction products formed in wells A-2, A-3, and A-4.
9. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. For each well, determine whether the Mn in MnO_4^- undergoes oxidation or reduction.
2. For each well, decide whether MnO_4^- acts as an oxidizing agent or a reducing agent.
3. Identify the other agent in the reaction, and indicate whether it is an oxidizing agent or a reducing agent.
4. Predict the effect of replacing $NaHSO_3$ with $NaHSO_4$. Test your prediction.
5. Complete the table involving reactions between $KMnO_4$ and $NaHSO_3$.

	Well	Change in Oxidation State of Mn	Change in Oxidation State of S	Electrons Gained by Mn	Electrons Lost by S
Reaction 1	A-2	7 → 6	4 → 6	1	2
Reaction 2					
Reaction 3					

6. In Reaction 1 of the table above, each MnO_4^- takes up one electron, but each HSO_3^- gives up two electrons. Account for this difference.
7. Predict the stoichiometric ratio of MnO_4^- to HSO_3^- in Reactions 2 and 3.
8. Balance each of the following skeleton reactions based upon the stoichiometric ratios developed in Question 7:

$$OH^- + MnO_4^- + HSO_3^- \rightarrow MnO_4^{2-} + HSO_4^- + H_2O$$
$$H^+ + MnO_4^- + HSO_3^- \rightarrow MnO_2 + HSO_4^- + H_2O$$
$$H^+ + MnO_4^- + HSO_3^- \rightarrow Mn^{2+} + HSO_4^- + H_2O$$

Implications and Applications

This activity involves observing a reaction in which oxidation states change. Color changes provide concrete evidence for reaction. The observed ratios suggest ways in which one can keep track of electrons gained or lost, and serve as the basis of a procedure for balancing oxidation-reduction equations.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Oxidation States of Manganese

Major Chemical Concept

The oxidation state of manganese in potassium permanganate is changed as the result of oxidation-reduction reactions.

Level

Introductory level for basic, general and honors students.

Expected Student Background

See *Related Concepts*. Students should have experience with small-scale equipment (8- or 12-well strips or 96-well plates; obtaining drops from Beral storage devices. See Frame A36950 and subsequent frames of the *Doing Chemistry* A-side videodisc for demonstrations of Beral pipet preparation.)

Time

One 50-minute class period.

Safety

Goggles and aprons worn. Wash spills of H_2SO_4 and NaOH immediately. Do not ingest chemicals. Dispose of manganese compounds by converting to Mn^{2+} and flushing down the drain with large amounts of water or converting to MnO_2 and disposing in accord with local regulations. Although not a part of the activity, teachers should know that mixtures of potassium permanganate and concentrated sulfuric acid lead to the formation of explosive manganese oxide, Mn_2O_7.

Materials (For 24 students working in pairs)

0.01 M $KMnO_4$, potassium permanganate, 250 mL (0.4 g $KMnO_4$ per 250 mL solution)
3 M H_2SO_4, sulfuric acid, 250 mL (42 mL conc. H_2SO_4 per 250 mL solution—add acid to water)
2.0 M NaOH, sodium hydroxide, 250 mL (20 g NaOH per 250 mL solution)
0.01 M $NaHSO_3$, sodium bisulfite, 250 mL (0.26 g $NaHSO_3$ per 250 mL solution)
0.01 M $NaHSO_4$, sodium bisulfate, 250 mL (0.30 g $NaHSO_4$ per 250 mL solution)
Twelve 24-well plates (or small test-tubes)
Beral pipets (or droppers)
White (bond) paper
Round toothpicks

Advance Preparation

Prepare the 0.01 M $KMnO_4$ and $NaHSO_3$ fresh. Within 48 hours of the laboratory activity, transfer these solutions to Beral storage pipets. (The $KMnO_4$ and $NaHSO_3$ solutions are reactive and change upon storage.)

Pre-Laboratory Discussion

Suggest that atoms of one element may have different oxidation states depending upon the compound. Stress the importance of making detailed observations, noting colors, textures, bubbles, *etc.* The number of drops of $NaHSO_3$ may need to be increased to obtain the proper colors.

Question 4 of the Data Analysis Section asks students to predict the effect of replacing $NaHSO_3$ with $NaHSO_4$. The easiest way to test this prediction is to repeat the procedure replacing $NaHSO_3$ with $NaHSO_4$. If students confuse these two substances, they will get unexplainable results. Therefore, during pre-laboratory discussion, you may want to point out this potential misidentification problem.

Teacher-Student Interaction

No special monitoring is required. Help students focus upon detailed observations and recording notes as you walk around the laboratory.

Anticipated Student Results

The alkaline solution turns green and then slowly turns murky. A brown precipitate forms in the well with no added acid or base. The acid solution turns colorless.

Answers to Data Analysis and Concept Development

1. MnO_4^- is reduced in each well.
2. MnO_4^- is the oxidizing agent in each well.
3. HSO_3^- is the reducing agent in each well and is oxidized.
4. In HSO_4^-, the sulfur is already oxidized to the +6 oxidation state. This ion is not a reducing agent.
5.

	Well	**Change in Oxidation State of Mn**	**Change in Oxidation State of S**	**Electrons Gained by Mn**	**Electrons Lost by S**
Reaction 1	A-2	7 → 6	4 → 6	1	2
Reaction 2	A-3	7 → 4	4 → 6	3	2
Reaction 3	A-4	7 → 2	4 → 6	5	2

6. Two permanganate ions react for each bisulfite ion that reacts.

$$2\,OH^-(aq) + 2\,MnO_4^-(aq) + HSO_3^-(aq) \rightarrow 2\,MnO_4^{2-}(aq) + HSO_4^-(aq) + H_2O(l)$$

   This may be thought of as each manganese atom lowering its oxidation number from +7 to +6 as the result of picking up one electron, while, at the same time one sulfur atom releases 2 electrons thereby raising its oxidation number from +4 to +6.

7. The ratios are 2 MnO_4^- / 1 HSO_3^- in A-2; 2 MnO_4^- / 3 HSO_3^- in A-3; and 2 MnO_4^- / 5 HSO_3^- in A-4.
8. A-2: $2OH^-(aq) + 2MnO_4^-(aq) + HSO_3^-(aq) \rightarrow 2MnO_4^{2-}(aq) + HSO_4^-(aq) + H_2O(l)$

   A-3: $2H^+(aq) + 2MnO_4^-(aq) + 3HSO_3^-(aq) \rightarrow 2MnO_2(s) + 3HSO_4^-(aq) + H_2O(l)$

   A-4: $6H^+(aq) + 2MnO_4^-(aq) + 5HSO_3^-(aq) \rightarrow 2Mn^{2+}(aq) + 5HSO_4^-(aq) + 3H_2O(l)$

Post-Laboratory Activities

Discuss the observations, working toward a consistent set of class observations. Discuss the assignment of oxidation states.

Assessing Laboratory Learning

Ask specific questions about the substances and reactions studied. Also pose open-ended items such as:

1. $NaHSO_3$ is frequently used to remove brown rust stains. Suggest a reason for this use. What must you know about the rust stains to make this suggestion consistent with principles? *Iron(III) oxide and hydroxide are much less soluble in acid solution than iron(II) hydroxide and oxide. The* $NaHSO_3$ *reduces iron(III) to iron(II).*
2. Before safety concerns (high toxicity) caused it to be dropped from common use, laboratory activities were performed with ammonium vanadate, $(NH_4)_3VO_4$. Acidic yellow solutions of this compound were shaken with solid zinc metal, Zn. The solutions underwent spectacular color changes. Suggest reasons that might account for these changes, and describe the kind of information you would need to have or experiments you would want to perform to confirm your hypotheses. *Zinc metal reduces vanadium(V) through vanadium(IV) and vanadium(III) intermediates to vanadium(II).*

$$4H^+ + 2VO_2^+ + Zn \rightarrow 2VO^{2+} + Zn^{2+} + 2H_2O$$

$$4H^+ + 2VO^{2+} + Zn \rightarrow 2V^{3+} + Zn^{2+} + 2H_2O$$

$$2V^{3+} + Zn \rightarrow 2V^{2+} + Zn^{2+}$$

Each species has a unique, characteristic color in acid aqueous solution. One way to confirm this series of reactions would be to measure the potential of the solution during the reduction process and relate potential changes to color changes. The reverse titration is often performed. Solutions of vanadium(II) are titrated with cerium(IV) to produce, in stepwise fashion, vanadium(III), vanadium(IV), and finally vanadium(V).

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: The Silver Mirror Reaction

Reference: The Woodrow Wilson microscale version of this lesson was developed by Fen Lewis, Strongsville High School, 7701 Beach Road, Wadsworth, OH 44281, and was based upon Kemp and Marwin. (1981). Journal of Chemical Education, 58, 655 .

Introduction

Everyday mirrors that we take for granted are manufactured using redox reactions in which silver compounds are reduced to silver metal. Because the silver metal becomes bonded to the glass so that tarnishing of the silver is slow, a shiny 'silver mirror' lasts for months or years, depending on methods of preservation.

Purpose

An ordinary silver mirror is produced using typical laboratory materials.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Nitric acid is corrosive. Silver nitrate causes stains.
3. Handle these substances with caution. Wash spills immediately with large amounts of water. Hold your thumb on the stopper while shaking.
4. Always mix the solutions fresh and dispose of them immediately after use with large amounts of water. The materials may form explosive silver fulminate, $Ag_2C_2N_2O_2$, on standing. *Never* premix the reagents.
5. Dispose of all materials as your teacher directs.

Procedure

1. Thoroughly clean with laboratory detergent the inside of a test-tube given to you for this laboratory activity. A scrupulously clean test-tube is essential for successful completion of this activity.
2. Rinse the test-tube with distilled water. Taking care to keep the acid off your skin, rinse again with 1/4 test-tube full of 6 M nitric acid, followed by a rinse with distilled water. Discard under hood.
3. Rinse again with acetone, then with distilled water.
4. Add the following amounts of solutions in the *exact* order listed:

   16 Drops stabilized honey solution (5%). Roll the tube to wet it with this solution before adding silver nitrate.
   8 Drops 8.0% silver nitrate solution
   8 Drops 12% ammonium nitrate solution
   16 Drops 10% sodium hydroxide solution

5. Quickly stopper the test-tube with a cork and shake it while holding the cork in place. The inside surface of the test-tube should be wetted for a good coating. Continue shaking the test-tube for about 3 min. Observe the changes in appearance for 5 min.
6. Wash the solution down the drain with 1 L water. This is an important safety precaution to prevent the possible formation of an explosive mixture (after standing many hours or days). Rinse the mirrored test-tube gently but thoroughly with distilled water. Allow the tube to air dry.
7. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. What is the visible product of the reaction?
2. Did the silver ions gain or lose electrons in the reduction process?
3. An oxidation reaction must take place along with a reduction reaction. Honey (actually the sugars in honey) is the partner for the reaction of silver ions in this activity. What happened to the sugar molecules?

Implications and Applications

The production of mirrors is the direct application illustrated by this activity. Essentially all metallurgical processes involve redox reactions. When metals are obtained from ores or are purified, the process involves reducing the element to its metallic state or oxidizing some impurity to a compound that can be separated from the remaining metal.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: The Silver Mirror Reaction

Major Chemical Concept

The controlled reduction (*i.e.*, slow reduction) of silver ion leads to the formation of metallic silver.

Reactions:

Formation of the reagent (Tollen's Reagent):

$$2AgNO_3 + 2NaOH \rightarrow Ag_2O + H_2O + 2NaNO_3$$

$$Ag_2O + 4NH_3 + H_2O \rightarrow \underset{\text{Tollen's reagent}}{2Ag(NH_3)_2OH}$$

Reduction of the reagent:

$$RCHO + 2Ag(NH_3)_2OH \rightarrow RCOO(NH_4) + 2Ag + 3NH_3 + H_2O$$

Level

Introductory level for basic, general, and honors students.

Expected Student Background

See *Related Concepts and Skills.*

Time

One 50-min class period

Safety

1. Always mix the chemicals fresh and dispose of them immediately after use with large amounts of water. The chemicals may form explosive silver fulminate, $Ag_2C_2N_2O_2$, on standing. *Never* premix the chemicals.
2. Sometimes teachers use acetone to clean glassware thereby removing grease-containing materials from the glass surface. Also, some teachers recover used glassware by removing silver with concentrated nitric acid. Mixtures of these two materials, acetone and concentrated nitric acid, are explosive.

Materials (For 24 students working in pairs)

24 Test-tubes, 15- x 120-mm, and corks to fit them.
5% Honey, 200 mL (10 g honey in 100 mL water. Add 1.2 g tartaric acid, boil, then cool the solution. Add 20 mL ethanol and dilute with water to 200 mL)
8% Silver nitrate, $AgNO_3$, 100 mL (8.0 g $AgNO_3$ per 100 mL solution)
12% Ammonium nitrate, NH_4NO_3, 100 mL (12.0 g NH_4NO_3 per 100 mL solution)
10% Sodium hydroxide, NaOH, 200 mL (20 g NaOH per 200 mL solution)
6 M Nitric acid, HNO_3 (add 76 mL concentrated nitric acid to 80 mL distilled water and dilute to 200 mL)
Acetone, 200 mL

Advance Preparation

Obtain the honey from a food store.

Pre-Laboratory Discussion

Focus the pre-laboratory discussion on practical applications of redox reactions and controlling the rate of a reaction. Combustion reactions, for example, are usually very fast redox reactions while rusting and tarnishing are slow.

Teacher-Student Interaction

Pay particular attention to safety. Several dangerous substances are in use. As you check individual students, check that the test-tubes have been properly cleaned. Dirty test-tubes are a main cause for failure and/or poor results. You may also need to coach the shaking process to make certain that adequate shaking occurs without making it too vigorous. (You may ask students to place corked tubes inside plastic zip-closure bags for shaking. Have them hold the cork in place while shaking even if a zip-closure bag is used. This strategy is helpful when the material being shaken presents either toxicity or stain-producing hazards.)

Anticipated Student Results

A laboratory full of beautiful silver mirrors will appear after about 35 min of work.

Answers to Data Analysis and Concept Development

1. Metallic silver
2. Silver ions gained electrons in reduction
3. Sugar molecules are oxidized.

Post-Laboratory Discussion

Discuss the original definition of the term reduction (reducing a metal to an element from a compound or ore) in the context of mirror formation. Point out that there are organic and biochemical reducing agents such as glucose, as well as inorganic reagents such as $NaHSO_3$, and that the energy for living systems is derived from biochemical redox reactions. Discuss the relative quantity of energy released from this reaction as compared with, say, the energy released in combustion of methane in a burner.

Assessing Laboratory Learning

Ask specific questions about the chemicals and reactions. Consider asking a safety question about the storage of Tollen's reagent.

DEMONSTRATIONS

Demonstration 1: Reduction of Copper(II) Oxide

(Modified from Experiment A45 of Doing Chemistry*)*

Description

Black copper(II) oxide is reduced to form copper metal (with its characteristic appearance) using methane gas. The reaction is clean and stoichiometric.

Materials

Burner
Hose—with bent glass tubing inserted into rubber stopper or cork for clamping
Test-tube Pyrex™, 25- x 200-mm (16- x 100-mm)
Wire to suspend tube for weighing
Copper(II) oxide, finely powdered, 1 g
Source of methane gas
2 Single-buret clamp
2 Support stands
Balance
May need Y-tubes and clamps to provide sufficient gas outlets
Safety shield

Laboratory Hints

An interesting variation is to use a tightly fitting two-holed rubber stopper for the tube. Insert a long glass tube through one hole so that it comes within 1 cm of the bottom of the test-tube and extends 5 cm to the other side of the stopper. Insert a 10-cm length into the other hole so that 3.5 cm of tubing protrudes on each side of the stopper. Connect rubber tubing from the gas source to the long glass tube. Connect rubber tubing from the burner to the short glass tube. Turn on the gas. The gas first flows over the oxide, and then flows out through the burner. Wait 15 seconds. Light the gas. When the reaction is complete, remove the burner, and allow the reaction test-tube to cool while the gas burns in the burner. Y-tubes or T-tubes and suitable clamps may be needed to provide enough gas outlets in laboratories where these are not available.

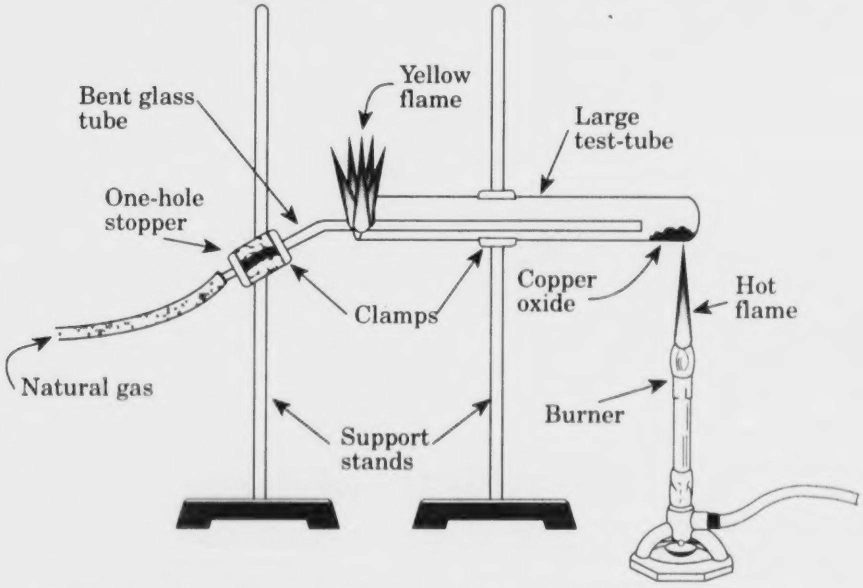


Figure 1. Apparatus for the reduction of copper(II) oxide.

Discussion

A possible error source is partial reduction of black CuO to red Cu_2O, which is often mistaken for Cu.

Safety

Natural gas is flammable and potentially explosive. Hot glassware can cause burns. Bending glassware and inserting glass tubing through rubber stoppers can lead to severe cuts.

Precautions

1. Use a safety shield.
2. Ignite escaping gas after a 15-second wait. Be careful handling hot glassware; use fingers to sense whether glassware has cooled to room temperature before grasping. Do not use plastic coated clamps.
3. To insert tubes into stoppers, use fire-polished glass (*i.e.*, glass heated to remove sharp edges.) Use a lubricant such as glycerin on the glass and the stopper. Wrap the glass and the stopper in a cloth towel. Slowly, with minimal pressure, twist the glass into the stopper.

Disposal

Reuse the same test-tube from year to year. The solid may be disposed of with ordinary solid waste.

Procedure

1. Set up the apparatus behind a safety shield.
2. Prepare a wire from which to hang a test-tube. Wrap a loop of wire around the top of a 16-mm test-tube near the lip. Shape the other end of the wire into a loop. This loop should be about 3-4 cm from the tube. Determine the mass of a test-tube by hanging it from the loop of wire attached to the balance hook.
3. Weigh a clean, dry test-tube. Record the mass.
4. Add about 1 g copper(II) oxide with a metal scoop so that as little as possible of the powder gets on the side of the test-tube.
5. Reweigh the test-tube and powder. Record the mass.
6. Insert the glass tube connected to the methane gas supply carefully so that it does not touch the powder. You may want to hold it during the demonstration or clamp it to a support stand. The glass tube should extend far into the test-tube, but not touch either the solid in the test-tube or the test-tube itself.
7. Turn on the source of methane gas for the glass tube inserted into the test-tube. After five seconds ignite the gas at the mouth of the tube. Adjust the methane flame at the mouth of the tube so that it is about 5 cm (2 in) high.
8. Light a burner. Heat the exterior of the test-tube to as high a temperature as possible with the burner for about 5 to 7 min or until copper color appears in the material in the test-tube.
9. Turn off the burner, but keep the methane flowing through the inside of the test-tube until the test-tube cools down (about 5 min). Once cooled, turn off the methane. (Because of the heat released at the top of the tube when the methane is burning, the tube will not cool to room temperature. Once the gas is turned off, wait until the tube cools to room temperature.)
10. Reweigh the test-tube and material remaining. Record the mass.

Sample Data Table

Mass empty tube	41.80 g
Mass tube plus oxide	43.00 g
Mass tube after reduction	42.78 g

Data Analysis

The mass of copper(II) oxide is: 43.00 g - 41.80 g = 1.20 g
The mass of solid remaining is: 42.78 g - 41.80 g = 0.98 g

If all remaining material were Cu, there would be:

0.98 g Cu x (1 mol Cu/ 63.54 g Cu) = 0.0154 mol Cu

If all the lost mass were due to oxygen, there would be:

1.20 g - 0.98 g = 0.22 g O, and
0.22 g O x (1 mol O/ 16.00 g O) = 0.0138 mol O

The mole ratio of Cu to O is:

0.0154 mol Cu/ 0.0138 mol O = 1.16

Questions

1. Write a balanced chemical equation to describe the reaction between methane and copper(II) oxide. *[$CH_4 + 3CuO \rightarrow CO + 2H_2O + 3Cu$]*
2. Use the sample data here. Assume that the formula of copper(II) oxide is CuO. For the sample size taken, predict the mass of copper that should remain after reaction.

   [1 mol Cu x (63.54 g Cu / 1 mol Cu) = 63.54 g Cu
   1 mol O x (16.00 g O / 1 mol O) = 16.00 g O
   63.54 g Cu + 16.00 g O = 79.54 g CuO
   1.20 g CuO x (63.54 g Cu / 79.54 g CuO) = 0.959 g Cu

   For CuO, the predicted mass of Cu remaining is 0.959 g.]
3. Compare the predicted mass of copper remaining with that actually observed. Suggest possible reasons for discrepancies. *[The observed mass is 0.98 g Cu, and the predicted mass is 0.96 g Cu. This difference can be due to weighing error or insufficient heating.]*

Demonstration 2: Sodium in Water

(Modified from Experiment C03 of Doing Chemistry*)*

Description

A small piece of metallic sodium is placed in a dish of water containing phenolphthalein on an overhead projector stage. The sodium reacts vigorously, melts, and sputters on the surface of the water as a sphere of continuously decreasing size until it has completely reacted. The moving sodium leaves a pink trail in the water.

Materials

Overhead projector and screen
Shield
Petri dish (or 1-L beaker)
Stainless metal spatula with small blade
Sodium metal (lump the size of a small, dried, split, green pea)
Water (ice)
1% Phenolphthalein (in 95% ethanol)
Dry paper towels
Distilled water
Tongs or forceps

Safety

Metallic sodium is very dangerous; it is a strong reducing agent and prone to highly exothermic redox reactions. Fires and burns are possible. Flying debris from explosions resulting from misuse of sodium may cause severe injury.

Precautions

1. Store the sodium in small amounts under kerosene or mineral oil (nonproton-donating, nonreactive solvents). Store only enough to use for this demonstration for the school year, and no more than 10 g. Wrap the bottle in a heavy-weight polyethylene bag. Store in a place where it will not get wet, or come in contact with acids or oxidizing agents. Check every three months to be sure that the storage procedures are being followed.
2. Do not touch sodium with your fingers.

3. Use only the amount recommended. The amount used is magnified by the overhead projector. Your students will almost certainly try to tempt you use a larger piece of sodium; *never* do this. Wear goggles; have all students wear goggles. Use cool water.

Disposal

1. Call for assistance (such as district science supervisor, local college chemistry department, local fire marshal) in destroying excess sodium. Sodium stored in small amounts according to the precautions noted presents no special danger or hazard to the school, students, custodians, or fire fighters.
2. After the reaction ceases, the reaction mixture may be disposed of safely in the sink.
3. Be certain that no small metallic sodium fragments remain on the towel before disposing of it. If fragments are noticed, use tongs to immerse the towel in a large plastic bucket filled with cold water.

Procedure

1. Set up an overhead projector. Place a safety shield between the projector and the students. Work from the student side of the projector to protect yourself. Wear goggles and have your students wear goggles. Place a Petri dish (or a 1-L beaker) on the projector stage. Pour some cool water into the dish, and add 1-2 drops of 1% phenolphthalein in ethanol. Cool with ice if necessary. Check for visibility by students, and adjust as needed.
2. Remove the bottle of sodium metal from its storage area. Place on a safe, secure surface. Unscrew the cap. Use a dry spatula to "stab" a small chunk of the metal. Lift this chunk out of the bottle with tongs or forceps and place on a dry paper towel.
3. Use the spatula to cut a very small piece of sodium from the chunk. The piece should have a volume no larger than a rice grain. Note the surface of the freshly cut sodium. Bring this to your students' attention. Have them come close to view the surface, but under no circumstance let them touch the sodium. Do not pass the sodium around the room.
4. Return the chunk of metal to the storage oil as soon as a small piece has been cut off. Tighten the cap.
5. Place the small piece of sodium on the tip of the spatula. Transfer to the Petri dish. Note evidence for reaction.
6. Wear eye protection while disposing of the contents of the Petri dish at the sink. Flush with moderate amounts of water.

Demonstration 3: The Blue Bottle Reaction

(Modified from Experiment A06 of Doing Chemistry*)*

Description

One solution is divided between two containers (flasks) such that one container is nearly full and the other nearly empty. Shaking the filled container does not seem to produce a dramatic change, but shaking the partially filled container does. Pouring the liquids from the filled container to the empty one and repeating the procedure leads to the same result—only the nearly empty container gives a dramatic change.

Materials

2 Florence flasks, 1-L
Beaker, 2-L
Stirring rod, 40-cm
Glucose (or fructose), 10 g
Sodium hydroxide, NaOH, 10 g
1% methylene blue, in dropping bottle (dissolve 1 g methylene blue in 100 mL distilled water)
Distilled water

Laboratory Hints

1. Sucrose (table sugar) is not a suitable substitute for glucose.
2. Make a fresh glucose solution on the day of the laboratory activity.
3. This laboratory activity can be performed by students. Students may shake the solutions in stoppered test-tubes. Emphasize the safety precautions. (Remember, every student in the class will be shaking a strongly alkaline solution in a glass container!)

Safety

The solution is alkaline (about 0.2 M NaOH) and a skin irritant with prolonged exposure. Methylene blue will stain clothing.

Precautions

Wear eye protection. Wash off spills. Wear old clothing or laboratory aprons. Check that the safety shower and eye wash fountain are in working condition.

Disposal

Neutralize the sodium hydroxide solution with vinegar; dispose of at the sink.

Procedure

1. Place 1500 mL distilled water in a 2-L beaker. Add to the water 10 g glucose, 10 g sodium hydroxide, and sufficient drops of 1% methylene blue to color the mixture blue.
2. Fill a 1-L Florence flask with this solution. Pour the remaining solution into a second 1-L Florence flask.
3. Swirl each flask vigorously, and note results.
4. Pour from the filled flask into the partially filled flask and swirl the contents of both flasks vigorously. Note observations.
5. Pour once again, shake once again, and note results.

Discussion

Ask students to account for their observations. Once discussion slows down, assuming no one picks up on the role of air, suggest that one of the reactants, which they did not consciously add to the system, might be invisible. Ask, "What reactant should be added to the list?" Generate suggestions for testing their hypotheses.

Background

1. Oxygen oxidizes glucose well in a strongly alkaline solution. Methylene blue is also subject to oxidation. The oxidized form is blue; the reduced form (leuco methylene blue) is colorless. When oxygen is present, methylene blue is blue. When the solution is permitted to stand without shaking, the glucose reduces the methylene blue to the colorless form.

2. Shaking the partially filled flask causes oxygen gas to dissolve and to oxidize the methylene blue. When there is no source of fresh air—a filled flask, or one with very little surface—the blue does not form. You will probably be able to see a bluish color at the top surface of the solution. This careful observation provides support for the mechanism suggested to account for these color changes. In a sealed flask, once all of the oxygen gas is used up, there is no longer a changing of the color. Care must be taken to use fresh solutions.
3. Readily oxidized sugars may replace the glucose. Several redox dyes may replace methylene blue.
4. This demonstration can be used as an effective model for lakes in which there has been eutrophication (enrichment). When the environment is very rich, algae blooms and excessive weeds lead to oxidation of organic matter and subsequent oxygen depletion.

Small-scale

Some teachers implement this as an activity in which students operate at the test-tube scale. Students must wear eye protection at all times during the microscale activity. One way to present the laboratory activity is to provide one filled and one half-filled 4-mL screw cap vial and ask whether "the two solutions are the same."

Demonstration 4: Hydrogen Peroxide as an Oxidizing and Reducing Agent

Materials

2% Potassium iodide, KI (2 g KI per 100 mL solution)
0.02 M Potassium permanganate, $KMnO_4$ (0.3 g $KMnO_4$ per 100 mL solution)
3% Hydrogen peroxide, H_2O_2 (commercially available)
3 M Sulfuric acid, H_2SO_4 (Add 1 mL concentrated H_2SO_4 slowly to 5 mL distilled water.)

Safety

All of the substances are toxic. All except KI are corrosive.

Procedure

1. Place 100 mL KI solution in a large beaker. Add 6 drops 3 M H_2SO_4. Add 3% H_2O_2 until a brown color is produced.
2. Place 100 mL $KMnO_4$ solution in a large beaker. Add 6 drops 3 M H_2SO_4. Add 3% H_2O_2 until the purple color disappears.
3. This procedure works very well on an overhead projector.

Background

When H_2O_2 acts as an oxidizing agent, water is produced.

$$H_2O_2(aq) + 2I^-(aq) + 2H^+(aq) \rightarrow 2H_2O(l) + I_2(aq)$$

The iodine reacts with I^- to produce a colored complex, I_3^-.

$$I_2(aq) + I^-(aq) \rightarrow I_3^-(aq)$$

When H_2O_2 acts as a reducing agent, oxygen is produced:

$$2MnO_4^-(aq) + 5H_2O_2(aq) + 6H^+(aq) \rightarrow 2Mn^{2+}(aq) + 5O_2(g) + 8H_2O(l)$$

Disposal

Treat the Mn^{2+} solutions with $KMnO_4$ in neutral solution. Add sodium bisulfite. Filter the solid MnO_2 and hold for disposal with other solid heavy metal wastes.

Demonstration 5: Displacement of Silver by Copper

(Modified from Experiment A44 of Doing Chemistry*)*

Description

Solid silver metal can be obtained from a reaction in which metallic copper is the reducing agent and is oxidized. Copper is usually considered an easily reduced element. This chemical contrast together with the very attractive reaction (beautiful silver crystals, blue colored solution) make for an excellent demonstration.

Materials

0.1 M Silver nitrate, $AgNO_3$ (1.7 g per 100 mL solution)
Copper wire, 30 cm,
Graduated cylinder, 100-mL

Safety

Silver nitrate is toxic and causes dark stains.

Precautions

Handle silver nitrate solutions cautiously. Wipe up spills immediately.

Procedure

Place 100 mL 0.1 M silver nitrate in a 100-mL graduated cylinder. Coil the copper wire into a spring-like shape with a coil diameter smaller than the internal cylinder diameter. Place the copper coil in the silver nitrate solution. Note evidence for reaction. Allow to stand for several hours.

Reaction

$$2Ag^{+}(aq) + Cu(s) \rightarrow 2Ag(s) + Cu^{2+}(aq)$$

Demonstration 6: Potassium Permanganate and Glycerol

(Chemical Demonstrations, B. Z. Shakhashiri, Volume 1, p. 83*; shown on* Redox *videodisc, A27)*

Description

Solid potassium permanganate reacts with certain alcohols. The reaction with glycerol is dramatic because it starts slowly. As the reaction proceeds and the mixture heats up, it becomes extremely vigorous. The reaction rate depends upon the size of the potassium permanganate particles. Finely ground potassium permanganate reacts more quickly than does a sample with large crystals. In a side-by-side comparison, placing similar amounts of glycerol on two separate piles of potassium permanganate with different crystal sizes, the effect of surface area on reaction rate can be demonstrated.

Materials

Potassium permanganate, $KMnO_4$, 5 g
Glycerol, 2 mL
Dropping device (Beral pipet)
Evaporating dish, 90-mm
Spatula
(For reaction rate activity, need mortar and pestle and two of each listed item.)

Safety

Potassium permanganate is toxic and causes burns and stains. The reaction is extremely exothermic and can cause burns and uncontrolled fires.

Precautions

Handle the potassium permanganate cautiously. Perform the reaction away from combustible materials. Keep a loaded fire extinguisher nearby.

Procedure

Place 5 g $KMnO_4$ in an evaporating dish. Use a spatula to make a small indentation in the pile of solid $KMnO_4$. Add 1 mL glycerol to the indentation in the center of the pile of solid. Observe for several minutes. *(NOTE: Varying particle size of $KMnO_4$ provides an interesting lesson on surface area* vs. *reactivity.)*

Reactions

Several reactions are possible, such as:

$$14\,KMnO_4(s) + 4\,C_3H_5(OH)_3(l) \rightarrow 7\,K_2CO_3(s) + 7\,Mn_2O_3(s) + 5\,CO_2(g) + 16\,H_2O(l)$$

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. Can an atom be oxidized in the absence of oxygen? *[Yes. Chemically speaking, oxidation is not linked to oxygen. Oxidation is a loss of electrons.]*
2. Why aren't oxidizing agents oxidized? *[Oxidizing agents, the agents that bring about oxidation, bring about a loss of electrons by taking in electrons. They are reduced because they act as oxidizing agents.]*
3. Why aren't reducing agents reduced? *[Reducing agents bring about reduction by giving up electrons. To give up electrons is to become oxidized.]*
4. Using a number line, explain why gaining electrons does not make the resulting particle more positive. *[Gaining electrons moves to the left on the number line reducing the oxidation number. When the number line concept is used, gaining electrons is treated algebraically—gaining something negative.]*
5. How can an atom gain electrons and be "reduced?" *[This is a quirk in terms. In everyday usage, "reduction" implies a loss. By defining reduction as a gain in electrons, chemists create confusion with the everyday use of the word reduction.]*
6. Explain how it is possible for a substance to act as an oxidizing agent in one case and a reducing agent in another case. *[If a substance gives electrons in one reaction but takes electrons in another, it is a reducing agent in the first case, and an oxidizing agent in the second.]*

   $$H_2O_2 + MnO_4^- \rightarrow O_2 + Mn^{2+} \text{ (Not balanced)}$$
   (The oxygen in hydrogen peroxide undergoes oxidation)

   $$H_2O_2 + I^- \rightarrow I_3^- + H_2O \text{ (Not balanced)}$$
   (The oxygen in hydrogen peroxide undergoes reduction)
7. Explain why all reactions cannot be classified as redox reactions. *[Some reactions are acid base reactions. Other reactions involve forming solids by 'replacements.' These reactions do not involve changes in oxidation numbers of atoms.]*
8. How can a redox reaction be identified? *[A redox reaction is always identified by identifying a change in oxidation number.]*
9. How can one infer that bodily exercise increases oxidation of stored food? *[Exercise leads to weight loss. The rate of respiration during exercise increases, so that oxygen consumption increases and carbon dioxide production increases.]*

Counterintuitive Examples

Magnesium Burning in Carbon Dioxide *(shown on Redox videodisc, A08)*

Description

Magnesium is a strong enough reducing agent to reduce CO_2 to carbon. Carbon dioxide is regarded as something that extinguishes fires, but here supports the burning of magnesium. When something burns in an atmosphere of carbon dioxide, it is a counterintuitive event.

Materials

Dry ice (or sodium bicarbonate, marble chips, and vinegar)
Beaker, large
Wood splint
Magnesium, Mg, ribbon
Tongs
Burner and ignition source

Hazards

Burning magnesium causes burns. Magnesium ribbon can cause cuts. Dry ice can cause burns. Effervescing acid/carbonate mixtures can spray corrosive liquids.

Precautions

Wear goggles. Handle burning magnesium with tongs. Do not touch burning materials.

Procedure

1. Place dry ice in a beaker and wait about 10 min. (Alternatively, place about 30 g sodium bicarbonate and 15 g marble chips in a large beaker and carefully add 200 mL vinegar.) Ignite a wood splint in a burner flame, and plunge the burning splint into the gas at the top of the beaker. Ignite a piece of magnesium ribbon in a burner flame, and repeat.
2. The wood splint stops burning. The magnesium ribbon continues to burn, and a black solid (carbon) is produced.

$$2Mg + CO_2 \rightarrow 2MgO + C$$

In this reaction, the carbon dioxide is the *oxidizing agent* and becomes reduced.

The Can Ripper

An aluminum soda can is weakened as the result of a redox reaction. Select a 'sturdy' looking student from the class. Give the student an untreated pop can; ask the student to 'tear the can in half.' Then, select a student of small stature whom one would not suspect to be capable of such a feat. Give the student a treated can, and request that the same act be performed. A properly treated can is as easy to tear as is a piece of aluminum foil used for cooking.

Materials

Aluminum soda can
Copper(II) chloride, $CuCl_2$, 2 g
Tap water

Hazards

Unsuspecting persons may accidently drink from the can during the treatment. Do not allow this. The edges of the torn can may cause cuts.

Precautions

Keep the can stored in a place where students cannot accidentally use it during the treatment. Wear gloves when 'ripping' the can.

Procedure

Score the *inside* of an aluminum soda can by inserting a triangular file into the can opening and scratching a line inside around the can (about 5 cm from the bottom). Place 1-2 g copper chloride in the can, and add enough water to cover the scoring. Swirl to dissolve the copper(II) chloride. Wait about 3-5 min. (Experiment with the time to avoid "eat through" of the can.) Pour the solution into a heavy metal waste container. Rinse the can. Give a student garden gloves, and ask the student to tear the can. (Chloride ion is needed for this procedure to work effectively. Copper(II) sulfate is not a suitable substitute for copper(II) chloride unless a chloride source such as NaCl or HCl also is added.)

Reaction

$$2\text{Al(s)} + 3\text{Cu}^{2+}\text{(aq)} \rightarrow 2\text{Al}^{3+}\text{(aq)} + 3\text{Cu(s)}$$

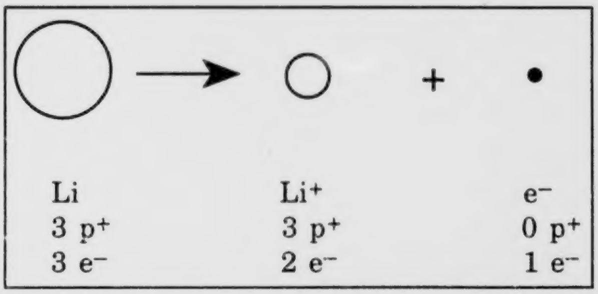
The copper(II) ions oxidize the aluminum metal.

Metaphors, Analogies, and Mnemonics

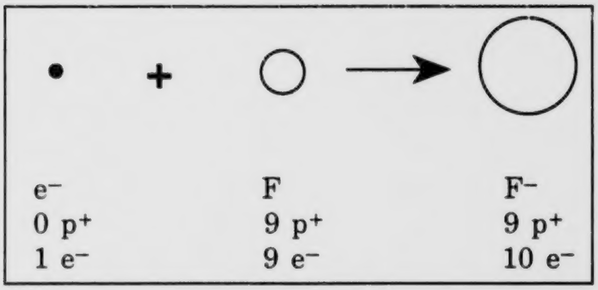
1. Two excellent memory devices are available to help students keep redox terms in mind. Give a hearty growl as you walk into the classroom to introduce "*LEO* the lion says *GER*" (Loss of Electrons is Oxidation...Gain of Electrons is Reduction). Or have one of your art students make a colorful poster to place on the chemistry room wall showing LEO in full majesty saying GER.
2. A second equally effective memory device is "*OIL RIG*" (Oxidation Is Loss...Reduction Is Gain). Again, a poster of an oil well spouting with the above statement included makes a great classroom decoration.

Pictures in the Mind

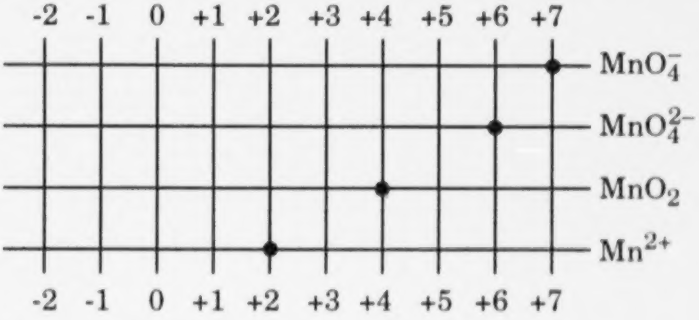
1. Draw a picture of an atom being oxidized.



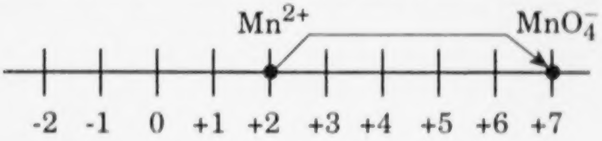
2. Draw a picture of an atom being reduced.



3. Using number lines, illustrate what happens to an atom that is reduced.

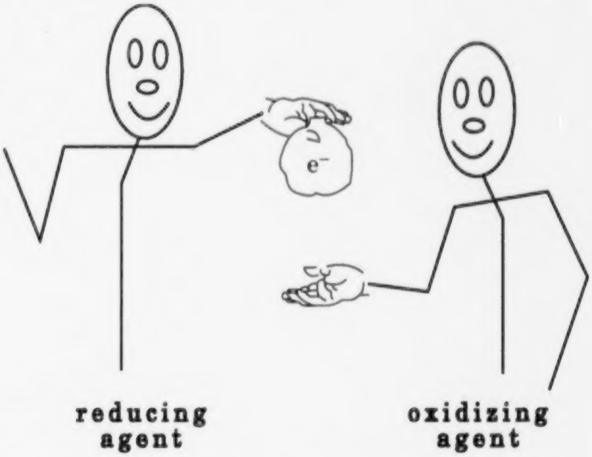

-2 -1 0 +1 +2 +3 +4 +5 +6 +7
MnO_4^-
MnO_4^{2-}
MnO_2
Mn^{2+}
-2 -1 0 +1 +2 +3 +4 +5 +6 +7

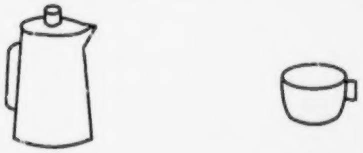
4. Illustrate movement on a number line to indicate oxidation.


Mn^{2+}
MnO_4^-
-2 -1 0 +1 +2 +3 +4 +5 +6 +7

5. Show a picture of an oxidizing agent at work.

6. Show a picture of a reducing agent at work.


e⁻
reducing agent
oxidizing agent



coffee pot
filling *agent*
is emptied

coffee cup
emptying *agent*
is filled

TIPS FOR THE TEACHER

Language of Chemistry

combustion chemical reaction accompanied by heat and light, usually associated with oxidations using oxygen as the oxidizing agent.

corrosion reaction of metals forming compounds, particularly when exposed to air and water (to corrode is to eat away at or destroy slowly); metallic corrosion involves oxidation of the metal.

electron negatively charged subatomic particle with low mass.

oxidation reaction in which oxidation number increases (electrons are lost to another species).

oxidation number number used in keeping track of the change in distribution of electrons during reaction; it is assigned on the basis of several rules such as the oxidation number of an atom in its elemental form is zero.

oxidation state (see oxidation number)

oxidize to increase in oxidation number; to lose electrons to another species.

oxidizing agent chemical entity that brings about oxidation and becomes reduced; an entity that gains or picks up electrons during chemical reaction.

redox acronym used to describe reactions in which oxidation numbers change.

reduce to decrease in oxidation number; to gain electrons from another species.

reducing agent chemical entity that brings about reduction and becomes oxidized; an entity that loses electrons during chemical reaction.

reduction reaction in which oxidation number decreases (electrons are acquired from another species).

respiration biochemical reactions that provide energy through the metabolism or breakdown of biochemicals.

Oxidizing and Reducing Agents

After your students are comfortable with the concepts of oxidation and reduction (with the help of *LEO* says *GER* or *OIL RIG*), the concept of oxidizing agents and reducing agents should be introduced. Ask the class to define the term agent. Connect your discussion with previous experiences in earth science where students may have learned about "agents of erosion." Ask them to list the agents of erosion. When they tell you wind, water, *etc.*, ask them why these are the "agents" of erosion. They will most certainly state that they are agents because they "cause" erosion.

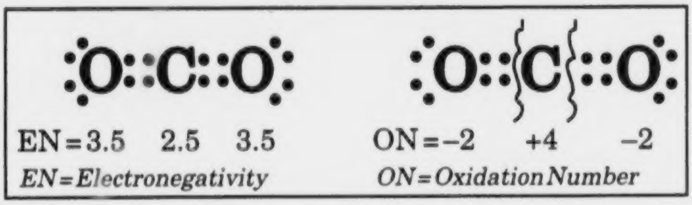
Tie this concept of an agent as a "causer" to the concept of oxidizing/reducing agents. An oxidizing agent is a "causer" of oxidation. Since oxidation is loss of electrons (*OIL*), an oxidizing agent *causes loss of electrons*. It should follow that if the oxidizing agent causes the loss of electrons by another substance, it is *gaining electrons*. The next logical step in the discussion is to help students see that if the oxidizing agent is gaining electrons, it is being reduced (*RIG*).

Correspondingly, a reducing agent must be a "causer" of reduction. Since reduction is gain of electrons (*RIG*), a reducing agent causes another substance to gain electrons. The only way it can cause gain of electrons is to lose electrons. If the reducing agent is losing electrons, it is being oxidized, since oxidation is loss (*OIL*).

To bring the discussion to closure, ask students to explain why a substance cannot be an oxidizing agent and, in the same reaction, be oxidized. If they understand the concepts involved, they will quickly see why oxidizing agents are not oxidized and why reducing agents are not reduced.

Assigning Oxidation Numbers

Oxidation number can be thought of as the charge a bonded atom would have if all bonding electrons were assigned arbitrarily to the more electronegative atom.



Oxidation Numbers are Assigned Using Several Arbitrary Rules

1. The oxidation number of an atom in its elemental state is 0. [S_8, P_4, O_2]
2. The oxidation number of an atom in a monatomic ion is equal to the charge on the ion. [$Fe^{3+} = +3$, $O^{2-} = -2$, $N^{3-} = -3$]
3. The oxidation number of fluorine atoms in a compound is –1. [CF_4, $F = -1$]
4. The oxidation number of alkali metal atoms in compounds is +1. [$Na^+ = +1$; $K^+ = +1$]
5. The oxidation number of alkaline earth metal atoms in compounds is +2. [$Ca^{2+} = +2$; $Mg^{2+} = +2$]
6. The oxidation number of aluminum in compounds is +3.
7. When halogen elements are in compounds with less electronegative elements, halogen atoms have oxidation numbers of –1. [Cl in NaCl = –1; Cl in $PCl_3 = -1$; Cl in HCl = –1]
8. Hydrogen has an oxidation number of +1 with more electronegative elements. [H in HI = +1]
9. Oxygen usually has an oxidation number of –2 in compounds.
10. The sum of oxidation numbers of all atoms in a neutral molecule is 0. The sum of oxidation numbers of all atoms in an ion is equal to the charge of the ion.

Other Rules *(Not always taught in high school chemistry)*

1. The oxidation number of an atom in a (homonuclear) polyatomic ion is equal to the charge on the ion divided by the number of atoms in the ion. [$Hg_2^{2+} = +1$, $I_3^- = -1/3$]
2. Hydrogen has an oxidation number of –1 with less electronegative elements. [H in NaH = –1]
3. When combined with fluorine, oxygen has a positive oxidation number [O in $OF_2 = +2$]; in peroxides, oxygen has an oxidation number of –1 [O in $H_2O_2 = -1$; O in $BaO_2 = -1$]; oxygen has an oxidation number of –1/2 in superoxides [O in $KO_2 = -1/2$].

Using the Algebraic Number Line with Redox Equations

Students may have difficulty determining the number of electrons lost or gained in a reaction. They generally have difficulty deciding which substance involved is oxidized and which is reduced. The number line concept they may have learned in elementary algebra can be utilized to good advantage here.

A move to the right on a number line leads to a more positive number; a move to the left leads to a more negative number. In redox chemistry, since electrons with negative charges are being "moved around," the concept is best illustrated by stating both moves in terms of negatives; that is, more positive can be viewed as being less negative. [An interesting aside is that this same process operates in electronics where, for example, a potential of +5 V can be considered to be negative with respect to a potential of +10 V. Since it is 5 V less positive, it is, in effect, 5 V more negative. When both points where voltmeter leads are to be connected are above ground potential (positive), electronic experimenters soon learn to connect the black (negative) lead from the voltmeter to the less positive (more negative) point.]

For a redox example, imagine that Fe^{3+} ions are converted to Fe^{2+} ions. Have students draw a simple number line to illustrate this oxidation number change. The number line move is from +3 to +2 (to the left). The oxidation number is now less positive (more negative) by one charge. The ion must have gained one electron to become one charge more negative. The number is *reduced* in magnitude by a move to the left. This fact will assist students in associating a move to the left on the algebraic number line with chemical reduction.

If I^{7+} is converted to I^-, there is a move of eight units to the left on the number line. The number is obviously less positive (more negative). Again, it is apparent that 8 electrons have been "gained" by the ion. Iodine has been "reduced" in both algebraic and chemical senses.

If one reverses the process so I^- is converted to I^{7+}, then the move is eight to the right on the number line. The resulting ion is eight charges more positive (less negative) than I^-. The only way an ion can become less negative is to lose electrons. The original ion must have been oxidized (*OIL*) in the chemical reaction. Encourage your students to use the algebraic number line concept to aid their understanding of redox concepts.

Pattern Recognition

1. More active (lower electronegativity; lower left-hand portion of Periodic Table) metals replace (reduce) less active (higher electronegativity) metals from a compound.
2. More active (higher electronegativity) nonmetals replace (oxidize) less active (lower electronegativity) nonmetals from a compound.
3. Metals reduce nonmetals. Nonmetals oxidize metals.
4. As an atom's oxidation state moves left on the number line, the species is being reduced. As its oxidation state moves to the right on the number line, the species is being oxidized.

Common Student Misconceptions

1. **"All the bonds formed are ionic, as shown by the use of oxidation numbers."**

   Oxidation numbers are assigned to atoms in molecules or ions that are described as being made up from essentially covalent bonds.
2. **"Oxidation number represents the actual charge of the atom in a molecule."**

   This is not so. For example, the iodine atom in IO_4^- is not thought of as a 7+ cation surrounded by four 2– anions, but instead as having electron pairs shared between the iodine and each oxygen atom.
3. **"All reactions are redox reactions."**

   Acid base reactions are not redox reactions; precipitation reactions such as $BaSO_4(s)$ from $Ba^{2+}(aq)$ and $SO_4^{2-}(aq)$ are not redox reactions.
4. **"Oxidation always involves oxygen."**

   Oxidation involves a loss of electrons. Oxygen atoms are good at gaining electrons, but so are fluorine atoms, chlorine atoms, *etc.*

5. **"Electrons are lost when something is 'reduced.'"**

   Electrons are 'lost' when captured by certain radioactive nuclei or when they react in negative electron/positive electron pairs. Electrons are not lost (*i.e.*, are not consumed) in redox reactions. Instead, they may be thought of as moving from one species (the reducing agent) to another (the oxidizing agent).

6. **"A half reaction can occur by itself, *e.g.*, oxidation without reduction."**

   In real systems, reduction and oxidation are paired. The electrons required to bring about a reduction come from an oxidation. If an oxidation were to occur by itself without accompanying reduction, the electrons would still need to go someplace—and extra electrons, in large numbers, repel one another with enormous force.

7. **"All oxidation numbers must be integers."**

   Oxidation number rules sometimes lead to fractional oxidation numbers. For example, the oxidation number of oxygen in potassium superoxide (KO_2) is –1/2 (see *Other Rules* in *Language of Chemistry*).

Problem Solving

Students should:

1. Identify redox reactions from given equations.
2. Identify what is oxidized and what is reduced in a redox reaction.
3. Identify oxidizing and reducing agents in a redox equation.
4. Use rules to assign oxidation numbers to atoms in molecules, ions, and salts.
5. Balance simple redox equations when given the redox rules while keeping track of electron loss and gain.

Samples

1. $Zn(s) + 2H^+(aq) \rightarrow H_2(g) + Zn^{2+}(aq)$

   Zn, the reducing agent, is oxidized. The oxidation number of Zn changes from 0 to +2. H^+, the oxidizing agent, is reduced. Its oxidation number changes from +1 to 0. The oxidation numbers are assigned as 0 for elements and are set equal to the charges for monatomic ions.

2. $2SO_2 + O_2 \rightarrow 2SO_3$

   SO_2, the reducing agent, is oxidized. The oxidation number of S changes from +4 to +6. O_2, the oxidizing agent, is reduced. Its oxidation number changes from 0 to –2. The oxidation number of O in O_2 is 0 (it is an element). The oxidation number of S in SO_2 is +4. Oxygen is assigned –2 in compounds. Two times –2 is –4. Since the molecules have no net charge, this –4 must be balanced by the sulfur, so sulfur's oxidation number must be +4. Similarly, the oxidation number of S in SO_3 is +6.

3. $2Ag^+(aq) + Cu(s) \rightarrow 2Ag(s) + Cu^{2+}(aq)$

   $Ag^+(aq)$, the oxidizing agent, is reduced. Cu(s), the reducing agent, is oxidized. One mole of Cu(s) is oxidized by two moles of $Ag^+(aq)$.

Additional Examples

For each reaction depicted, identify the oxidizing agent, the reducing agent, the species oxidized, and the species reduced. Indicate whether the equation is balanced. Give the oxidation numbers of all atoms other than hydrogen and oxygen.

1. $2Cu^{2+}(aq) + Fe(s) \rightarrow Cu(s) + Fe^{2+}(aq)$

   [Equation unbalanced; balanced equation is:

   $$2Cu^{2+}(aq) + Fe(s) \rightarrow 2Cu(s) + Fe^{2+}(aq)$$

   $Cu^{2+}(aq)$ is oxidizing agent and is reduced; $Fe(s)$ is reducing agent and is oxidized. Oxidation numbers: Cu in $Cu^{2+}(aq)$, +2; Fe in $Fe^{2+}(aq)$, +2; Cu in $Cu(s)$, 0; Fe in $Fe(s)$, 0.]

2. $6H^{+}(aq) + 2MnO_4^{-}(aq) + 5HSO_3^{-}(aq) \rightarrow 2Mn^{2+}(aq) + 5HSO_4^{-}(aq) + 3H_2O(l)$

   [Equation balanced. MnO_4^{-} is oxidizing agent and is reduced; HSO_3^{-} is reducing agent and is oxidized. Oxidation numbers: Mn in MnO_4^{-}, +7; S in HSO_3^{-}, +4; S in HSO_4^{-}, +6; Mn in Mn^{2+}, +2.]

3. $C_3H_7OH(l) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$

   [Equation unbalanced. Balanced equation is:

   $$2C_3H_7OH(l) + 9O_2(g) \rightarrow 6CO_2(g) + 8H_2O(l)$$

   O_2 is oxidizing agent and is reduced; C_3H_7OH is reducing agent and is oxidized. Oxidation numbers: C in C_3H_7OH, –2; C in CO_2, +4.]

HISTORY: ON THE HUMAN SIDE

1. **Luigi Galvani (1737 - 1798).** During the 1770s Galvani began research in the investigation of animal electricity in nerves and muscles. From 1780 to 1791, he did extensive research into the effects elicited by static electricity on properly prepared frogs. The preparation consisted of removing the spinal cord, crucial nerves and lower limbs of the frog. Galvani then conducted experiments examining the convulsive contractions of the muscles in the lower limbs. He soon discovered an unexpected result: the muscles contracted without electrical stimuli when the frog, placed on an iron plate, was prodded with brass hooks. Galvani had discovered the central phenomenon of galvanism and produced an example of a redox reaction; the production of electric current was possible through the contact of two different metals in a moist environment. However, Galvani did not recognize the redox reaction; instead he interpreted his results as evidence for a theory that animal muscles possess their own unique electrical fluid.

2. **Alessandro Guiseppe Antonio Anastasia Volta (1745 - 1827).** The first accurate analysis of the phenomenon observed by Galvani was provided by Volta in 1792. His research, using the whole frog, showed that the muscle discharge was the result of electricity produced by dissimilar metals coming in contact with one another through a moist environment. Volta demonstrated first that galvanic excitations arose from external electrical stimulation and, secondly, that metals could be ranked by their electromotive power. In 1793 he demonstrated that the electromotive power of a chain of dissimilar metals depended only on the nature of the two extreme links, those touching the moist conductor. Volta went on to measure the voltage of a variety of metals undergoing redox reactions. Volta debated these theories with Galvani during the early 1790s. Volta's explanations for the electromotive power of metals and their production of voltage eventually proved to be a more accurate description of the redox phenomenon than that of the galvanists.

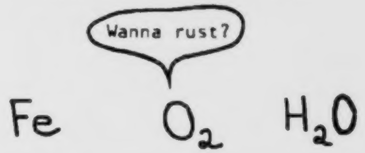
3. **Humphry Davy (1778 - 1829).** While Volta recognized that electricity was generated when two different metals were placed in solution, it was Humphry Davy who explained the chemical nature of this reaction. After reading Volta's work, Davy began working on electrolytic cells and demonstrated that a chemical reaction was generating the electricity.

4. **Hermann Walter Nernst (1864 - 1941).** Research on thermoelectricity provided the basis for Nernst's development of an electron theory of thermal and electrical conductivity. By examining solutions of various ionic concentrations separated by semipermeable partitions, Nernst soon demonstrated the relationship between ionic mobility, the diffusion coefficient and the electromotive force in concentration cells. As a result, he developed a mathematical equation for the electromotive force (E) in a galvanic process and a concentration change (C_1 to C_2): $E = \frac{RT}{NF} \ln \frac{C_1}{C_2}$

In this equation, R is the gas constant, T the absolute temperature, N the gram equivalents that have reacted, and F is Faraday's constant. Through this equation, a fundamental link between thermodynamics and electrochemical solution theory was established.

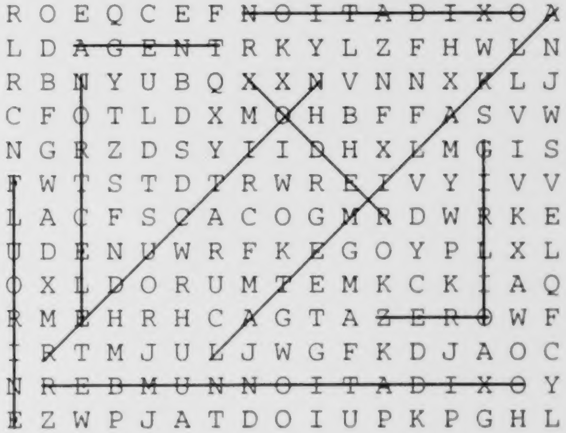
HUMOR: ON THE FUN SIDE

1. Student responses on exam:
   a. Reduction is burning wood giving off different elements and substances.
   b. Combustion is where there are a lot of greasy rags piled together and the air can't get through them. The air can go just so far and then it starts back. The air going in and coming out heats the rags and the rags get hot. Then they burst into flame.
2.



CHEM 13 NEWS, January 1983, p. 3

3. Word Search (see *Appendix* for master copy)



R	O	E	Q	C	E	F	~~N~~	~~O~~	~~I~~	~~T~~	~~A~~	~~D~~	~~I~~	~~X~~	~~O~~	A
L	D	~~A~~	~~G~~	~~E~~	~~N~~	~~T~~	R	K	Y	L	Z	F	H	W	L	N
R	B	N	Y	U	B	Q	X	X	N	V	N	N	X	K	L	J
C	F	O	T	L	D	X	M	O	H	B	F	F	A	S	V	W
N	G	R	Z	D	S	Y	I	I	D	H	X	L	M	G	I	S
F	W	T	S	T	D	T	R	W	R	E	I	V	Y	I	V	V
L	A	C	F	S	C	A	C	O	G	M	R	D	W	R	K	E
U	D	E	N	U	W	R	F	K	E	G	O	Y	P	L	X	L
O	X	L	D	O	R	U	M	T	E	M	K	C	K	I	A	Q
R	M	E	H	R	H	C	A	G	T	A	~~Z~~	~~E~~	~~R~~	~~O~~	W	F
I	R	T	M	J	U	L	J	W	G	F	K	D	J	A	O	C
N	~~R~~	~~E~~	~~B~~	~~M~~	~~U~~	~~N~~	~~N~~	~~O~~	~~I~~	~~T~~	~~A~~	~~D~~	~~I~~	~~X~~	~~O~~	Y
E	Z	W	P	J	A	T	D	O	I	U	P	K	P	G	H	L

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Loss of electrons.
2. Gain of electrons.
3. The entity that causes an event to occur.

4. Mnemonic used to remember which species gains and which loses electrons.
5. Charge a bonded atom would have if all bonding electrons were assigned to the more electronegative atom. (2 words)
6. Most highly electronegative element.
7. The sum of oxidation numbers of all atoms in a neutral molecule.
8. This family of atoms in compounds has an oxidation number of +1. (2 words)
9. Shortened form of writing oxidation-reduction.
10. Entity transferred in an oxidation-reduction reaction.

Answers: 1. OXIDATION 2. REDUCTION 3. AGENT 4. OILRIG 5. OXIDATION NUMBER 6. FLUORINE 7. ZERO 8. ALKALI METAL 9. REDOX 10. ELECTRON

4. See cartoons at end of module.

MEDIA

1. *The World of Chemistry* videotape "Number 15: The Busy Electron," can be used by chemistry teachers at all levels. World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.
2. The *World of Chemistry* (high school version) videotape "Number 17: Oxidation and Reduction," WINGS for Learning/SUNBURST, 101 Castleton Street, Pleasantville, NY 10570; (800) 321-7511; (914) 747-3310; (914) 747-4109 (FAX).
3. CHEM Study films/videos available from Ward's Natural Science Establishment, Inc., P.O. Box 92912, Rochester, NY 14692-9012; (800) 962-2660.
   a. Volume 24: *Vanadium, A Transition Element*
   b. Volume 20: *Nitric Acid*
4. *Doing Chemistry*, videodisc set available from the American Chemical Society, 1155-16th St., N.W., Washington, DC 20036; (202) 872-4382.
5. *Closeup on Chemistry*, video available from the American Chemical Society, 1155-16th St., N.W., Washington, DC 20036; (202) 872-4382.
6. Software published by *JCE: Software*, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
   a. *REDOX: A Tutorial on REDuction / OXidation Equations*, by Derek Davenport, Paul Groves and Dale Jensen. Vol. III A, No. 1 for the Apple II computer.
   b. *Oxidation Number Rules*, by James P. Birk. Vol. VI B, No. 1, for IBM PS/2, PC-compatible computers.
   c. *Frost Diagrams*, by James P. Birk and Heidi Hocker. Vol. IV C, No. 1, for the Apple Macintosh.

7. Videodiscs published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
   a. “Copper-Zinc Electrochemical Cell,” a chapter on *The World of Chemistry: Selected Demonstrations and Animations*: Disc I (double sided, 60 min.), Special Issue 3.
   b. “The Aufbau Principle,” “Making Sodium Chloride” and “Thermite Reaction, three chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc II (double sided, 60 min.), Special Issue 4.
   c. *The Periodic Table Videodisc* (single side, 30 min.), Special Issue 1.
   d. “Combustion of Methane,” “Oxidation of Alkanes, Alkenes and Alcohols,” and “Acetylene – Preparation and Reactions,” on the videodisc *Demonstrations in Organic Chemistry* (double sided, 60 min.), Special Issue 6.
8. *Redox*, by Helen Brooks and David Brooks. Videodisc distributed by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396; (608) 263-2837; (608) 262-0381 (FAX).

Links/Connections

WITHIN CHEMISTRY

1. **Electrochemistry**. Electrochemistry involves oxidation or reduction at the interface between a metallic conductor or semiconductor (an electrode) and an electrolyte. All electrochemical reactions are redox reactions.
2. **Stoichiometry**. Redox reactions are associated with changes of whole numbers of electrons (there are no fractions of electrons), so redox reactions involve substances in ratios of whole numbers as do other stoichiometric reactions.
3. **Periodicity**. The Periodic Table permits prediction of redox reactivity patterns of elements.
4. **Thermodynamics**. Most nonredox reactions studied (acid/base; lattice formation) have the energy changes accompanying the reactions limited in some special way by the nature of the reaction. (For example, aqueous solutions of acids and bases react to produce water. That main reaction, largely independent of particular acid and/or base sources, involves hydrogen ion and hydroxyl ion. The reaction [$H^+ + OH^- \rightarrow H_2O$] limits the quantity of energy released. This energy limit is much smaller than is the limit for energetic redox reactions such as the formation of water from hydrogen and oxygen.)

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

1. Biochemistry (majority of biochemical reactions are redox—*e.g.*, cell respiration.)
2. Food Science and Food Preservation
3. Geochemistry—*e.g.*, nitrogen cycle
4. Atmospheric Sciences—*e.g.*, photochemical smog (Nitrogen gas, N_2, is one of the most stable substances known, but when subjected to high temperatures as in internal combustion engines, it can be oxidized by atmospheric oxygen, O_2, to nitrogen(II) oxide, NO, and nitrogen(IV) oxide, NO_2. Nitrogen(IV) oxide is responsible for much of the "brown cloud" visible as smog in large urban areas.)
5. Civil Engineering—*e.g.*, corrosion prevention on infrastructures

TO THE CONTEMPORARY WORLD

Personal

1. Corrosion of the Statue of Liberty (see *ChemMatters*, April, 1985)
2. The sulfur from an egg smashed on an automobile may bring about car paint damage through redox reactions
3. Antioxidants in food
4. Decaying food
5. Polluted water caused by fossil fuel and high-sulfur coal burning. (Chemical oxidation of reduced carbon compounds—burning of coal, petroleum, wood—remains the major source of energy needed to maintain our lifestyles. Because much organic material contains sulfur as well as carbon, many air pollutants [SO_2, SO_3, CO, CO_2, NO, and NO_2] are the direct result of this energy production.)

Community

Field Trips

1. Local coal or gas fueled power plant
2. Electroplating factory
3. A metal producing factory (smelter, blast furnace, *etc.*)

Experts

1. Mining engineers
2. Chemists at electric power plants

Societal

1. **Decaying of metallic structures.** Corrosion, a major pathway of decay, requires replacement of existing structures made of steel.
2. **Photochemical smog and ozone layer.** Redox reactions lead to pollutants in the atmosphere. Some of these cause smog formation at ground level, while others destroy the ozone layer above.
3. **Greenhouse effect.** Although the oxygen quantity in the atmosphere is too extensive to be critically affected by human activities, the concentration of minor atmospheric gases can be changed drastically. Current concern about the greenhouse effect due to extensive oxidation of fossil fuels is a result of human activity. Even though it is a common gas we often mention when asked to name components of the atmosphere, CO_2 is present in only small amounts. The extent of fossil fuel burning is so great that the amount of CO_2 in the atmosphere is undergoing a significant percent increase. This gas prevents infrared radiation from leaving the earth and has the net effect of increasing the earth's temperature.
4. **Acid rain**. Sulfur in coal burns (oxidizes) to form sulfur dioxide. Sulfur dioxide is further oxidized to sulfuric acid in a series of steps including both oxygen and water. As the result of burning large amounts of coal, acid rain has played havoc with the ecological balance of many areas of the United States and Canada, especially in areas containing little or no limestone.
5. **Aging**. One of the principal causes of aging in organisms is attributed to effects of oxidation by 'radicals' formed when oxygen reacts in a biological system. Organisms have complex systems designed to prevent and to overcome the effects oxidative aging. On a different level, the freshness of nonliving food materials is often extended by adding substances called antioxidants. A simple example is the addition of vitamin C (ascorbic acid) when canning fruits to prevent browning (caused by air oxidation of the amino acid tyrosine and other phenolic materials followed by polymerization of the oxidation products).

References

Module developed by David Brooks, Robert Curtright, and James McGahan, the Nebraska team.

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Brooks, D. W. (Producer). (1989). *Doing chemistry.* [Videodiscs, computer program, and supporting written materials]. Washington, DC: American Chemical Society

Brooks, H. B., and Brooks, D. W. (Producers). *Redox.* [Videodisc, computer program, and written materials]. Washington, DC: American Chemical Society. (Distributed by Synaps, 334 South Cotner Blvd., Lincoln, NE 68510.)

Kemp, M. (1981). Tested demonstrations: Silver mirror. *Journal of Chemical Education, 58*, 655-656.

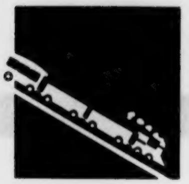
The Woodrow Wilson microscale version of Laboratory Activity 2: The Silver Mirror Reaction was developed by Fen Lewis, Strongville High School, 7701 Beach Road, Wadsworth, OH 44281, and was based upon this article.

Morgan, M. (1993, January). Ideas from the coast: Redox. *Connecting with Chemistry (Project SERAPHIM News)*, p. 3, 9.

Shakhashiri, B. Z. (1989). *Chemical demonstrations: A handbook for teachers of chemistry* (Vol. 3). Madison, WI: University of Wisconsin Press.

Summerlin, L. R., and Ealy, J. L. (1985). *Chemical demonstrations: A sourcebook for teachers*. Washington, DC: American Chemical Society

Appendix



- **Transparency Masters**
    1. Oxidation and Reduction—Redox
    2. Meet the Agents
    3. Word Search
- **Humor**

Oxidation and Reduction — Redox

Oxidation **is a loss of electrons:**

$$\mathrm{Zn} \longrightarrow \mathrm{Zn}^{2+} + 2e^-$$

Zinc is *oxidized.*

Reduction **is a gain of electrons:**

$$\mathrm{Cu}^{2+} + 2e^- \longrightarrow \mathrm{Cu}$$

Copper is *reduced.*

Oxidation and reduction always occur together:

$$\mathrm{Cu}^{2+} + \mathrm{Zn} \longrightarrow \mathrm{Cu} + \mathrm{Zn}^{2+}$$

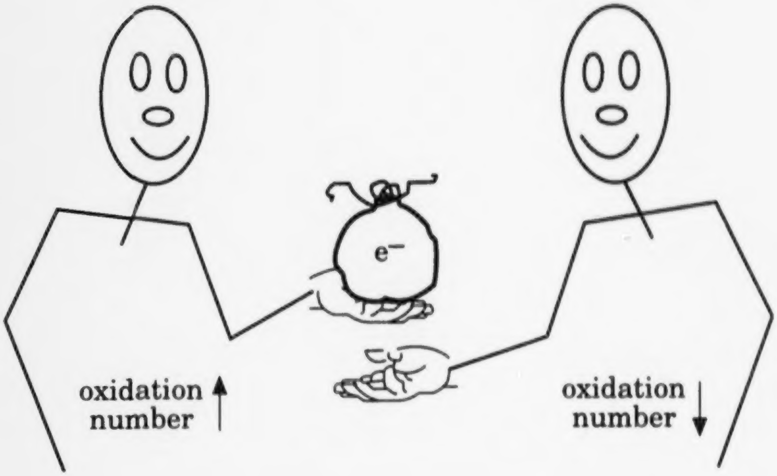
MEET THE **Agents**

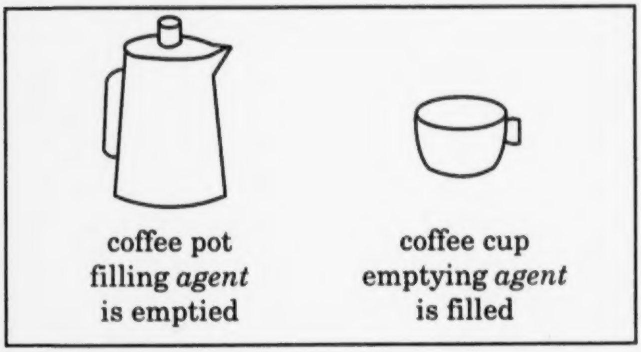
REDUCING

(loses electrons, becomes oxidized)

OXIDIZING

(gains electrons, becomes reduced)


e⁻
oxidation number ↑
oxidation number ↓



coffee pot
filling *agent*
is emptied

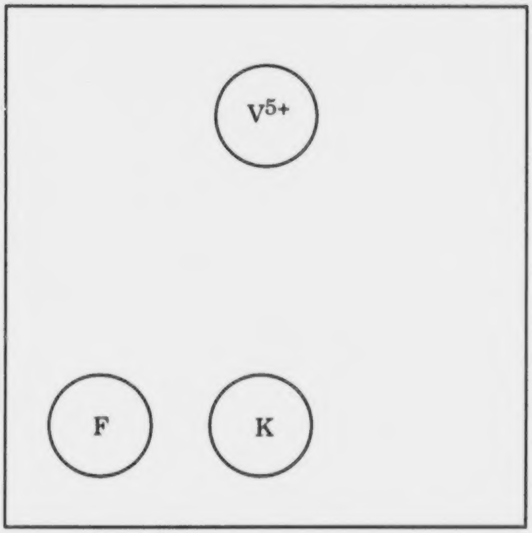
coffee cup
emptying *agent*
is filled

Word Search

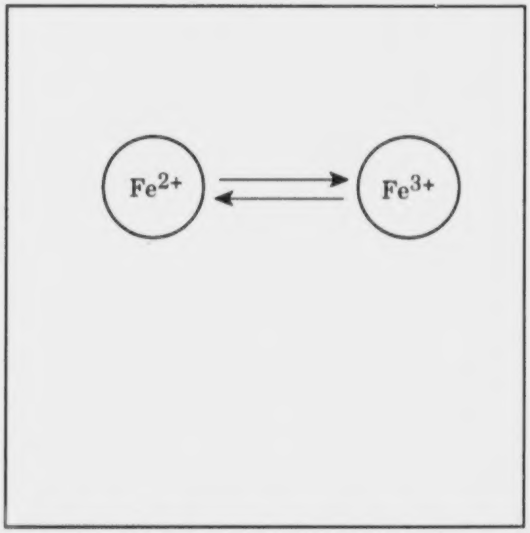
```
R O E Q C E F N O I T A D I X O A
L D A G E N T R K Y L Z F H W L N
R B N Y U B Q X X N V N N X K L J
C F O T L D X M O H B F F A S V W
N G R Z D S Y I I D H X L M G I S
F W T S T D T R W R E I V Y I V V
L A C F S C A C O G M R D W R K E
U D E N U W R F K E G O Y P L X L
O X L D O R U M T E M K C K I A Q
R M E H R H C A G T A Z E R O W F
I R T M J U L J W G F K D J A O C
N R E B M U N N O I T A D I X O Y
E Z W P J A T D O I U P K P G H L
```

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Loss of electrons.
2. Gain of electrons.
3. The entity that causes an event to occur.
4. Mnemonic used to remember which species gains and which loses electrons.
5. Charge a bonded atom would have if all bonding electrons were assigned to the more electronegative atom. (2 words)
6. Most highly electronegative element.
7. The sum of oxidation numbers of all atoms in a neutral molecule.
8. This family of atoms in compounds has an oxidation number of +1. (2 words)
9. Shortened form of writing oxidation-reduction.
10. Entity transferred in an oxidation-reduction reaction.



"Kay, the only problem with Van is you never know from day to day what state he'll be in!"



"Problem with Ferd is he can never seem to make up his mind."

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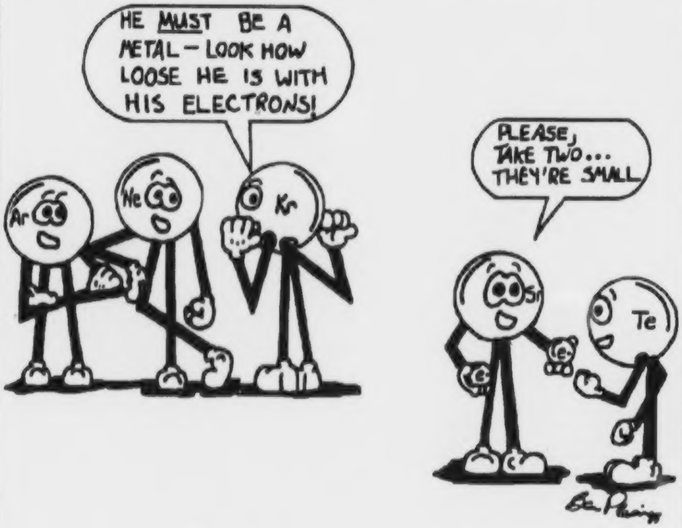
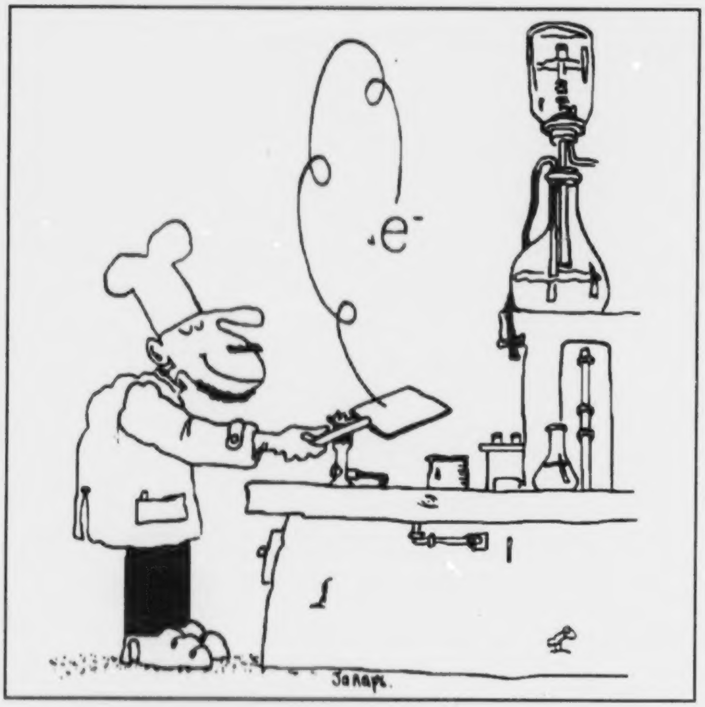
At parties, the noble gases didn't mix with the elements, but that didn't stop them from being terrible gossips.


HE MUST BE A METAL – LOOK HOW LOOSE HE IS WITH HIS ELECTRONS!
PLEASE, TAKE TWO... THEY'RE SMALL
Ar
Ne
Kr
Sr
Te

CHEM MATTERS, December 1988, p. 2. Reprinted with permission.



CHEM 13 NEWS, September 1988, p. 26. Reprinted with permission.

GEORGE, WHAT IS THAT?
A RED OX. THE BOOK SAID WE WERE GOING TO STUDY RED OX BEHAVIOR!
NO_WHAT I SAID WAS_WHAT I MEANT WAS_ THAT IS_

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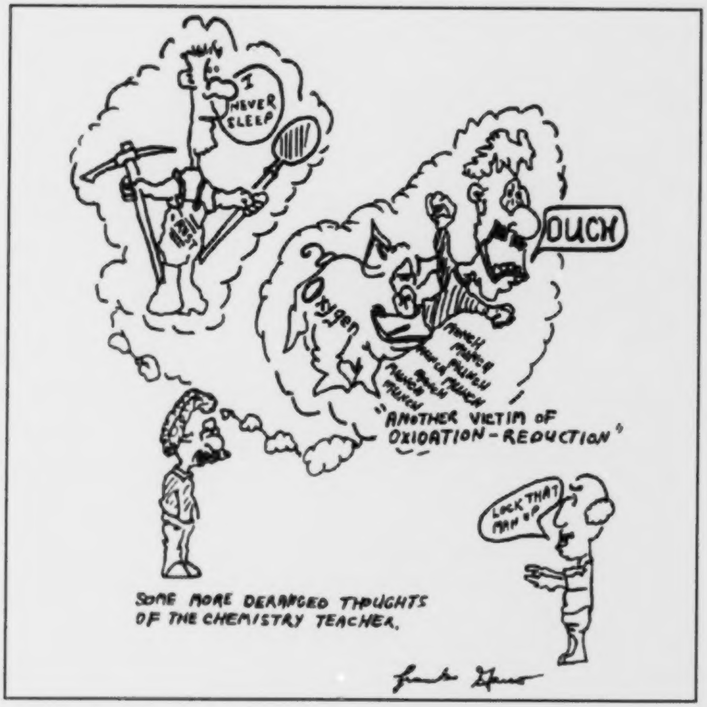

I NEVER SLEEP
OUCH
Oxygen
MUNCH MUNCH MUNCH MUNCH MUNCH MUNCH MUNCH MUNCH
"ANOTHER VICTIM OF OXIDATION-REDUCTION"
LOCK THAT MAN UP
SOME MORE DERANGED THOUGHTS OF THE CHEMISTRY TEACHER.

CHEM 13 NEWS, March 1981, p. 3. Reprinted with permission.

A SourceBook Module

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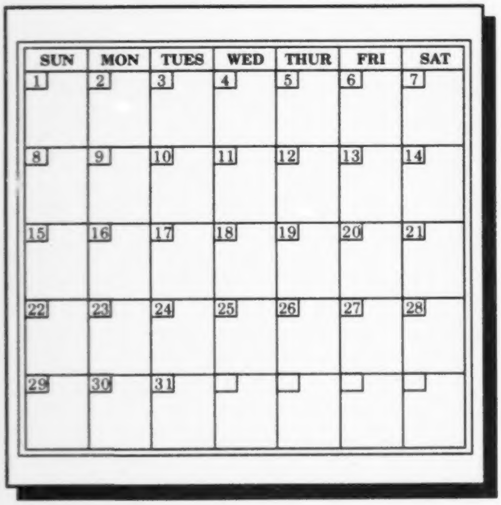
PERIODICITY


SUN MON TUES WED THUR FRI SAT
1 2 3 4 5 6 7
8 9 10 11 12 13 14
15 16 17 18 19 20 21
22 23 24 25 26 27 28
29 30 31

ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

Elements, the building blocks of everything in the universe, have a bewildering array of properties. It would be intimidating for chemists, and chemistry students to keep track of the properties of the elements and their compounds if it were not for some common characteristics allowing the elements to be grouped. Elements have periodic properties. This means if the elements are arranged in order of increasing atomic number, common properties occur at regular intervals (or cycles). The Periodic Table, found in most chemistry textbooks, is a two dimensional matrix or classification scheme based on this principle. In general, elements in the same column (or family) have similar properties. These properties vary as one goes across a row (or period) of the table. A knowledge of the format of the Periodic Table is an aid to predicting and remembering a vast amount of descriptive chemistry. This is one important reason for studying periodicity.

Two aspects in studying the Periodic Table are: (1) similarity of properties found in columns of chemical elements, and (2) trends in properties down a column or across a row of the table. Remember, however, that the concepts of periodicity are only a first approximation describing general characteristics. There are limits to periodicity. Not all properties show periodic trends, and not all families show similarities in properties. Some of the families are more representative for purposes of comparison. These are the first two and last three columns of the table. As one nears the center of the table, family characteristics become cloudy and unpredictable. This unpredictability occurs due to the complexity of electronic structure of these central elements. Consequently, linking electronic structure to chemical properties is a central emphasis of the study of periodicity. This is an important theoretical theme in chemistry and another important reason for studying periodicity.

As a classification scheme, the Periodic Table can be divided into a number of groups based on properties and electronic structure. Important ones are listed below with reference to Figure 1.

1. Hydrogen is considered as a group of one element because of its unique properties.
2. A diagonal stepped line (bold) divides the table into two unequal parts. Elements to the right of and above this line (as circled in the diagram) are **nonmetals**. Elements to the left of and below the line are **metals**. Many elements adjacent to this line (*e.g.*, Si and Ge) have both metallic and nonmetallic properties (such as semiconductivity) and are called **metalloids**.
3. The columns labelled p^1 through p^6 are called the *p*-block because the last valence electron enters a *p* orbital. This block contains families called the halogens (Column p^5) and the noble gases (Column p^6).
4. The columns labeled s^1 (the family named the alkali metals) and s^2 (the family named the alkaline earth metals) are called the *s*-block because the last valence electron enters an *s*-orbital. The circled Groups 3 and 4 on the diagram (the *s*- and *p*-blocks), when taken together, are called the **representative elements**.
5. This is the *d*-block or **transition elements**. The columns are labelled d^1 through d^{10} because the last valence electron enters a *d*-orbital.
6. This grouping is labeled the *f*-block. This block contains the **lanthanides** or rare earth elements (top row) and the **actinides** (bottom row). The columns are labeled f^1 through f^{14}. This block is sometimes grouped with transition elements.

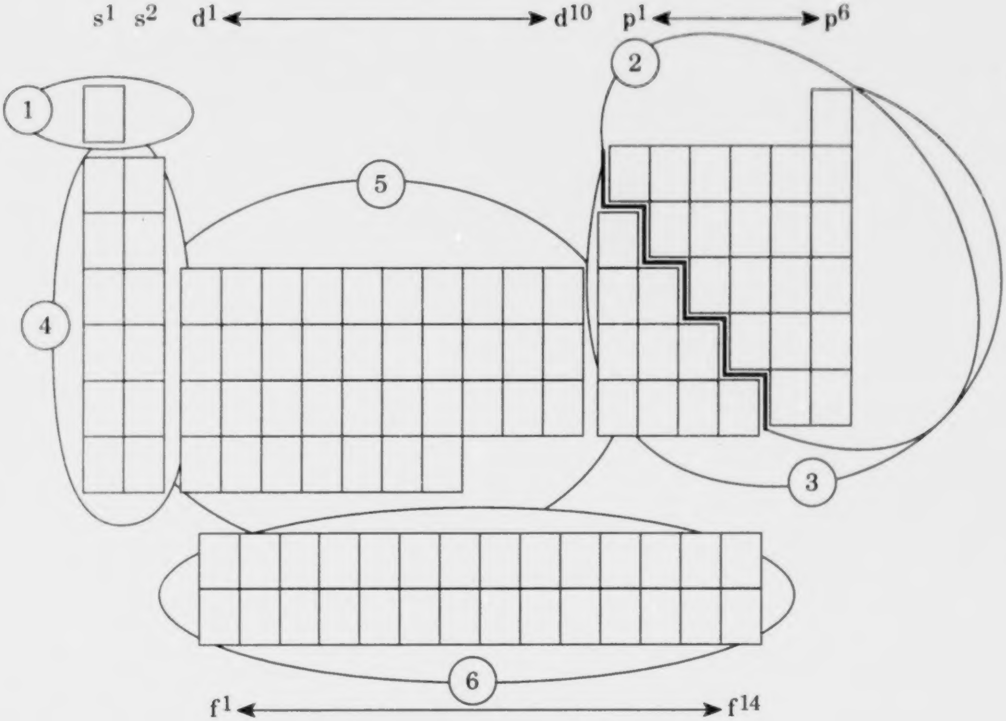


Figure 1. Divisions of the Periodic Table.

There are several ways of referring to the columns (groups or families) in the Periodic Table. In 1923, H.G. Deming published a popular chemistry text with a periodic table that eventually evolved into the long-form Periodic Table that we use today. In 1928, Merck and Company produced a convenient notebook form of the Table that was widely distributed in schools and eventually became the universally accepted Periodic Table in the United States. Deming numbered the columns of the representative elements with Roman numerals and an "A" designation. For example, he labeled the alkali metals Group IA, the alkaline earth Group IIA, the group headed by boron IIIA, *etc.* He labeled the transition elements with a "B" designation, so that the scandium group became IIIB, the titanium Group IVB, *etc.* This notation proved satisfactory to chemistry teachers for many decades because it was so easy to distinguish between the representative elements and the transition elements.

Meanwhile, European chemists adhered to a different designation for the Periodic Table columns. The Roman numerals and A and B designations were present but applied differently. European numbering simply went across the Table in sequence. Alkali metals were IA, alkaline earths were IIA, and the scandium group was IIIA. Continued numbering across the transition elements resulted in the copper group being designated IB, the zinc group IIB, and the boron group IIIB, *etc.* While Roman numerals continued to correspond to the outer electron structure of the elements within each group, the European scheme did not allow for the neat, clean distinction between the representative and transition elements. In 1970, the International Union of Pure and Applied Chemistry (IUPAC) recommended worldwide adoption of the European system. This recommendation spawned a great deal of confusion when some textbook publishers began to follow the IUPAC notation, and others continued to use the Deming notation.

In an attempt to bring some uniformity to the way in which the elemental groups were designated in the Periodic Table, the American Chemical Society (ACS) Division of Inorganic Chemistry Nomenclature Committee proposed a simple 1-18 (with a *d* designation for Groups 3-12) numbering scheme for the eighteen columns in the Periodic Table. This proposal was accepted as the official ACS form of the table in 1983. While the ACS solution eliminated ambiguity, the advantage of the group number corresponding to the number of valence electrons vanished.

American chemical educators continue to favor the Deming notation. Textbook publishers, in many instances, continue to cater to this preference. The result is a hodgepodge of Periodic Table numbering schemes. Many textbooks number the representative elements columns in the Periodic Table from 1 to 8, and don't number the transition element columns at all. The only way to really be safe is to cover all the bases and include all of the numbering schemes at the top of the Periodic Table, or to omit column numbering altogether. *SourceBook* editors have made the latter decision.

PLACE IN THE CURRICULUM

Periodicity and selected elements of the Periodic Table are usually taught in conjunction with and just after a consideration of the electronic structure of the atom in beginning chemistry courses. It is sometimes reintroduced in support of other subjects (especially those involving descriptive topics). Usually, representative elements are emphasized in an introductory module on periodicity because their properties are more predictable and the *s* and *p* valence electrons, which account for these properties, are easier to discuss theoretically. The properties of transition elements and the lanthanides and actinides depend in part on *d*- and *f*-electrons. Although a systematic study of these elements is usually not taught at the high school level, the *Transition Elements* module describes some of their properties.

CENTRAL CONCEPTS

1. Elements can be organized into a two dimensional matrix (periodic table) by arranging them in order of increasing atomic number.
2. Elements in a given column (family or group) of the Periodic Table show similar (related) properties.
3. Elements in columns (families or groups) and rows (periods) of the Periodic Table show trends in chemical and physical properties.
4. Elements in families and periods of the Periodic Table show trends in atomic properties (*e.g.*, atomic size).
5. Physical and chemical properties and an element's position in the Periodic Table can be related to electronic structure.

RELATED CONCEPTS

1. **Atomic Structure**
    a. Atoms are composed of a nucleus (containing protons and neutrons) and surrounding electrons. Electrons, though of little relative mass, are found in orbitals that occupy most of the atomic volume.
    b. Electrons occupy specific energy levels in an atom. Energy levels are quantized and are described by shells, subshells, and orbitals.
    c. Orbital occupancy of a multi-electron atom is given by its electron configuration.
2. **Bonding**
    a. Electrons at the highest occupied energy level in an atom in its ground state determine the atom's chemical properties and are called valence electrons.

   b. Atoms may form aggregates and gain stability by sharing an electron pair between them, thus forming a covalent bond.
   c. A metal atom and a nonmetal atom may form a stable structure by transferring of electrons from the metal atom to the nonmetal atom. This transfer forms oppositely charged particles called ions resulting in a structure held together by ionic bonds.
   d. Valence electrons are arranged around the kernels of atoms, forming molecules. Valence electrons can be represented as dots, yielding structures (Lewis-dot structures) consistent with the octet rule.

3. **Acids and Bases** (see *Acid-Base* module)
   a. Acids and bases are chemical species that exhibit different characteristic chemical properties.
   b. Acids are characterized conceptually as releasing H^+ ions and bases as releasing OH^- ions in water solution (Arrhenius Theory).
4. **Alkali Metals**
   a. Properties of alkali metals can be generalized in terms of chemical periodicity.
   b. Chemistry of alkali metals is based on the oxidation states 0 and +1.
5. **Halogens** (see *Halogens* module)
6. **Transition Elements** (see *Transition Elements* module)

RELATED SKILLS

1. **Chemical Skills**

   Write the name and symbol of common elements.
2. **Laboratory Skills**
   a. Use tests for common gases, especially $H_2(g)$.
   b. Use common acid/base tests.

PERFORMANCE OBJECTIVES

After completing their study of periodicity, students should be able to:

1. Explain trends and patterns in ionization energies of elements within families and periods of representative elements.
2. Explain trends and patterns in the sizes of elements within families and periods of representative elements.
3. Explain trends and patterns in the sizes of ions within families and periods of representative elements.
4. Predict relative sizes of atoms and ions in families and periods of representative elements not listed in the data tables given to students.
5. Explain the relationship between electron arrangement and location of elements in the Periodic Table.
6. Describe the nature of periods and groups (or families) in the Periodic Table.
7. Identify an element as an alkali metal, alkaline earth metal, halogen, or noble gas.
8. Identify an element as belonging to the *s*-, *p*-, *d*- or *f*-block in the Periodic Table.
9. State general and comparative reactivity trends for elements within a family of representative elements.
10. State general reactivity trends for representative elements within a given period.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Comparing the Reactivity of Aluminum, Calcium, Magnesium, and Sulfur with Water and Dilute Acid

Introduction

Reactivity, or the ease with which an element will react chemically, can be compared by observing reaction rates. You will experiment with several elements in combination with distilled water and then with dilute hydrochloric acid to determine the comparative reactivities of the elements. Your teacher may demonstrate the reactions of several other elements to add to those you investigated in the laboratory.

Purpose

To determine if any relationship exists between the placement of the selected elements on the Periodic Table and their chemical reactivities.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Use care in handling the solids. Pick up your samples with forceps to prevent possible chemical burns.
3. Be sure to cover beakers with watch glasses while observing the reactions.
4. Dispose of the reaction products and any remaining solid elements as your teacher directs.

Procedure

1. Pour about 5 mL of distilled water into an 18- x 150-mm test-tube. Test the water with litmus or pH paper before adding the solid, and before disposing of the test-tube contents. Add a piece of aluminum metal (Al) to the water.
2. Observe the system for 1-2 min, and record any evidence of chemical reaction. *NOTE: Heating a system sometimes speeds up a reaction. (Heating can be performed by placing your test-tube in a hot water bath.) If no change (or little change) is noticed after 1-2 min, heat the system for one minute, and record your observations.*
3. If there is evidence that a gas is being produced, test the gas with a burning splint (for H_2) or with a glowing splint (for O_2).
4. Dispose of the residue of this system as directed by your teacher.
5. Repeat Steps 1-3 separately using calcium, magnesium, and sulfur. Record your observations.
6. Perform similar tests on each of the elements using about 30 mL of dilute hydrochloric acid (HCl). Record observations of each system. Test the acid solution with litmus or pH paper before adding the solid, and before disposing of the contents.
7. Record your observations in the data tables.
8. Thoroughly wash your hands before leaving the laboratory.

Observations

Solids	Litmus or pH test before H_2O	Without heat	With heat	Test for H_2	Test for O_2	Litmus or pH test after reaction
Mg						
Ca						
Al						
S						

Solids	Litmus or pH test before HCl	Without heat	With heat	Test for H_2	Test for O_2	Test for Cl_2	Litmus or pH test after reaction
Mg							
Ca							
Al							
S							

Data Analysis and Concept Development

In interpreting the observations, the following chemical facts are needed:

> H_2 gas will burn rapidly in air; O_2 and Cl_2 gas will not. O_2 gas will support burning; a glowing splint thrust into O_2 will burst into flame. Cl_2 gas has a choking odor. A burning splint thrust into H_2 gas will cause an explosion.

1. List the elements that reacted with water in order of most to least vigorous.
2. List the elements that reacted with acid in order of most to least vigorous.
3. Compare the element order in the two lists.
4. Using the reactivity position of elements on your list, deduce any patterns you find between reactivity and element position on the Periodic Table.

Implications and Applications

1. Write balanced equations for all reactions observed between elements and water.
2. Write balanced equations for all reactions observed between elements and hydrochloric acid.
3. What chemical similarities are there between reactions of elements with water and elements with acid?
4. How would you explain the difference in reactivity between an element in water and an element in acid?
5. Using the results of your activities, how would you expect the reactivity of gallium (Element 31) to compare to the reactivity of aluminum, if both elements were placed in dilute hydrochloric acid? Explain your response.
6. State how you would expect strontium metal (Element 38) to react with water. Explain your reasoning.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Comparing the Reactivity of Aluminum, Calcium, Magnesium, and Sulfur with Water and Dilute Acid

Major Chemical Concept

Chemical reactivity of representative elements can be related to the element's position on the Periodic Table. This relationship is shown by reactions observed between several elements with water and dilute hydrochloric acid.

Level

General and Honors

Expected Student Background

1. Experience with observational evidence for chemical change.
2. Knowledge of atomic theory, electron structures of elements, and common ions.
3. Ability to write balanced equations given the reactants for reactions observed.
4. Knowledge of alkali metal reactivity trends, if possible. Do *Demonstration 1* (The Relative Reactions of the Alkali Metals) in combination with this activity.

Time

Two full class periods will be needed to perform the activity and discuss the results, if *Data Analysis and Concept Development* and *Implications and Applications* sections are assigned as homework. If this arrangement is not feasible, three to four class periods will be required.

Safety

1. Direct students to hold test-tubes with wire test-tube holders, and set test-tubes in a rack when not holding them.
2. If students wish to test the gas emitted for flammability, direct them to point the mouth of the test-tube away from themselves and other people.

Materials (For 24 students working in pairs)

12 Test-tubes, 18- x 150-mm
12 Wire test-tube holders
12 Forceps or tongs
12 Laboratory burners, ring stands, and wire gauze
12 Beakers, 250-mL
24 Pieces of aluminum shot (large pea size) or turnings
24 Strips of magnesium ribbon, 10-cm each (large pea size, when rolled tightly)
24 Pieces of lump sulfur (large pea size)
Calcium turnings (Ca oxidizes in air. Use of "old" Ca that has not been kept in an airtight container might not react at first until the protective oxide layer has been soaked through.)
Distilled water
6 M Hydrochloric acid, HCl, 2 L (Dilute 1000 mL conc HCl to 2000 mL total volume with distilled H_2O)
12 Test-tubes, 13- x 100-mm (for gas testing)
Red and blue litmus paper and/or pH paper

Advance Preparation

1. The metals and sulfur should be prepared and placed in labeled containers. Element samples should be of comparable size and conformation. You may wish to have students make a tight coil of the magnesium ribbon samples.
2. It is advisable to perform the activity prior to the class in which students perform the activity.

Pre-Laboratory Discussion

1. Review safety guidelines.
2. Although students are expected to decide what reactions occur and what products are obtained, suggest testing solutions with litmus paper and testing any gases produced for flammability. If so, review the method for collecting and testing gases, but do not tell them what gas to expect. Demonstrate tests for oxygen, hydrogen, and carbon dioxide gases. Refer to the *Identification of Gases* table in the *Appendix*.
3. Instruct students how to dispose of residues from the systems. Remind them to clean the beakers between tests.
4. *Demonstration 1* can be done prior to this activity. Students can then include Li, Na, and K in their set of elements for analysis.

Teacher-Student Interaction

Monitor student experimentation to ensure safety and correctness of procedure. Some students will be unsure if a reaction occurs in some systems. For example, gas bubbles may appear on the surface of solids, even though a strong reaction is not observed. In such cases suggest students tap the test-tube to loosen the gas from the solid and watch to see if more bubbles form. Students, however, should be allowed to determine whether or not a reaction occurred.

Anticipated Student Results

This chart summarizes the expected results. All the observed reactions involve evolution of hydrogen gas.

	H_2O	HCl
Al	N.R.	Reacts strongly
Ca	Reacts	Reacts violently
Mg	N.R.	Reacts very strongly
S	N.R.	N.R.

Answers to Data Analysis and Concept Development

1. Calcium. (No reaction occurred with the other three elements.)
2. Calcium, magnesium, aluminum, sulfur (no reaction)
3. Calcium is the most reactive in each set of data, but comparisons of the other elements can only be made by reacting them with acid.
4. Magnesium, aluminum, and sulfur are all in Period 3 of the table. Magnesium is more reactive than aluminum, and aluminum is more reactive than sulfur when combined with dilute acid. Magnesium and calcium are in the same family (column); data show calcium is more reactive than magnesium.

Answers to Implications and Applications

1. $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + H_2(g)$
   $Mg(s) + H_2O(l) \rightarrow$ N.R.
   $Al(s) + H_2O(l) \rightarrow$ N.R.
   $S(s) + H_2O(l) \rightarrow$ N.R.
2. $Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$
   $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
   $2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$
   $S(s) + HCl(aq) \rightarrow$ N.R.
3. When calcium reacts with water, there is a single displacement of hydrogen from water and the formation of hydrogen gas. This is an oxidation-reduction reaction; the calcium is oxidized and the hydrogen is reduced. The reaction is exothermic. Reactions of calcium, magnesium, and aluminum with acid share all the characteristics listed for calcium and water. One interesting difference between the calcium with water and calcium with acid reactions is that calcium reacts with water, forming a precipitate, but no precipitate forms in the reaction of calcium with acid. This difference is due to the low solubility of $Ca(OH)_2$, and the high solubility of $CaCl_2$ in water.
4. In all reactions hydrogen ion is reduced. Since the concentration of the hydrogen ion is much greater in acid than in water, and since increased concentration of reactants generally increases reaction rate, the reactivity of elements with acid is greater than it is with water. Sulfur reacted with neither water nor acid.
5. Gallium is listed directly below aluminum on the Periodic Table. Calcium was more reactive than magnesium, and it is located below magnesium. Since calcium, magnesium, and aluminum react by losing electrons (that is, oxidize) in these reactions, it would be reasonable to assume a similar type of reaction by gallium with acid. Calcium's two valence electrons are in the fourth energy level, whereas the two valence electrons of magnesium are in the third energy level. The higher reactivity of calcium can be explained, if one assumes that calcium's valence electrons are attracted less strongly than magnesium's because those of calcium are farther away from the attractive force of the nucleus. There are also more inner shell electrons to shield the nucleus. Given a similar situation in the comparison of electron arrangements of aluminum and gallium, one predicts that gallium will be more reactive than aluminum.
6. Strontium reacts more vigorously with water than calcium for the same reasons.

Post-Laboratory Discussion

A class discussion of the laboratory activity is essential. You may wish to base the discussion on student responses to the *Data Analysis and Concept Development* and the *Implications and Application* sections. Students' learning will be enhanced if they generate the answers to these questions, while you lead them with oral questions related to the way students are thinking about data obtained and questions asked in the *Student Instructions*.

Extensions

1. Perform *Demonstration 1* with the class, if this demonstration has not been performed already.
2. *Activity 2: Trends in Oxy-Hydrogen Compounds of Third Period Elements* is a valuable and reasonable extension to this laboratory activity.
3. *Activity 2* in the *Halogens* module involving the comparative reactivity of the halogen elements produces an important set of data that can be used to deduce periodic trends for nonmetallic elements.
4. Included in the *Group Activities* section of this module are data that can be graphed and the graphs used to deduce periodic patterns and trends for ionization energies, atomic sizes (radii), and ionic sizes (radii) for selected elements.

Assessing Laboratory Learning

1. Based on electron arrangements and common ions formed in chemical reactions, explain how the reactivity of sodium metal would compare to the reactivity of magnesium metal. *[Sodium and magnesium both have the third energy level as their valence level. Neutral atom electron arrangements are: Na (2-8-1), and Mg (2-8-2). Since Na has one electron removed to become Na^+ and Mg has two electrons removed to become Mg^{2+}, and the number of inner shell shielding electrons is the same for both, sodium should be more reactive. Further support for the argument is that the attraction of sodium's nucleus for electrons is caused by eleven protons, while magnesium has twelve protons in its nucleus.]*
2. What reactivity trend would you expect to find in the alkali elements, starting with lithium and going down the column? Explain your reasoning. *[The reactivity should be greater, as one goes from element to element down the column. Each of these elements reacts by having one electron removed from it. With each successive element, the distance of the valence electron from the nucleus increases by one energy level and electron shielding increases. These factors combine to make removal of the valence electron easier, and reactivity increases.]*
3. Based on your study of reactivity so far, how would you expect reactivities of lithium, beryllium, and boron to compare? Explain your reasoning. *[Lithium should be most reactive, followed by beryllium, and then boron. Lithium needs to have only one electron removed, while beryllium needs two electrons removed, and boron would have to have three electrons removed. These would be expected to follow the same pattern as their third period analogues, but should be less reactive because their valence electrons are closer to their nuclei and have less inner shell electron shielding. NOTE: This pattern is not observed as clearly, in actuality, as the pattern for the third period analogues.]*

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: Trends in Oxy-Hydrogen Compounds of Period 3 Elements

Introduction

Arrangement of elements in the Periodic Table can be used to establish interesting and important generalizations about chemical behavior. This laboratory activity will provide you with data leading to such a generalization.

Purpose

To use litmus and conductivity tests to determine properties of some oxy-hydrogen compounds of Period 3 elements.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Handle solutions with care. If you spill any solutions on your table or the floor, wipe up the spill and rinse the area with tap water.
3. If you should get any solution on your skin, flush the affected area with plenty of tap water.

Procedure

NOTE: Be sure to rinse your beaker well with tap water and then with distilled water after each test.

1. Obtain approximately 25 mL of one of the following solutions in a 50-mL beaker. Test the solution with litmus paper and with the conductivity apparatus. Record observations in the table below.

Solution	Litmus test	Conductivity test
NaOH		
MgO_2H_2		
$C_2O_2H_4$		
PO_4H_3		
SO_4H_2		

2. Repeat the procedure for each solution listed. Remember to wash your beaker between tests.
3. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

Chemical Facts needed for data analysis:

1. An acid is a compound that releases H^+ ions in aqueous solution. A base is a compound that releases OH^- ions in aqueous solution. An amphoteric compound is one that acts either as an acid or as a base depending on with what it reacts.
2. A water-soluble acid will turn blue litmus pink, will not affect pink litmus and will give a colorless solution when phenolphthalein is added. A water-soluble base will not affect blue litmus paper, will turn pink litmus blue and will give a red solution when phenolphthalein is added. A basic compound that is insoluble in water will often dissolve in an acidic (*e.g.*, HCl) solution. An acidic compound that is insoluble in water will often dissolve in a basic (*e.g.*, NaOH) solution.

From the results of litmus and conductivity tests, write the ions contained in each solution.

Solution	Ions present
NaOH	
MgO_2H_2	
$C_2O_2H_4$	
PO_4H_3	
SO_4H_2	

Implications and Applications

1. Make a general statement concerning the acid-base properties of the oxy-hydrogen compounds of the elements going across the Periodic Table.
2. What statement could you make from your data regarding the relative ability of Period 3 elements to form H^+ ions and OH^- ions?
3. Predict a litmus test, a conductivity test, and write the ions for a solution labeled ClO_4H.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: Trends in Oxy-Hydrogen Compounds of Period 3 Elements

Major Chemical Concept

A trend from basic to acidic properties can be observed in oxy-hydrogen compounds of representative elements in a period. This trend is observed as one goes from left to right across a period.

Level

General and Honors

Expected Student Background

1. Understanding which ions are responsible for acid or base characteristics of water solutions.
2. Knowledge of atomic theory.

Time

One class period will suffice for gathering the data. It is possible that some discussion of the data may occur in the first class period, although at least half of a second period will be required to discuss the laboratory activity.

Safety

Read the *Safety Considerations* in the *Student Version*. Use only 9-volt conductivity testers (not ones which employ 120-volt AC and bare wires). The 9-volt conductivity testers are easy to build, inexpensive to use and safer.

Materials (For 24 students working in pairs)

12 Beakers, 50-mL

Light bulb conductivity apparatus (See following note concerning construction of conductivity apparatus.)

Red and blue litmus paper

Sodium Hydroxide, NaOH, 500 mL (Dissolve two or three pellets of solid NaOH in 500 mL distilled water.)

Magnesium Hydroxide, $Mg(OH)_2$, 500 mL (Stir two spoonfuls of solid $Mg(OH)_2$ or milk of magnesia in 500 mL distilled water. Filter or allow the solid to settle and decant the supernatant solution.)

Acetic Acid, $HC_2H_3O_2$, 500 mL (Add 5-10 mL concentrated $HC_2H_3O_2$ to 500 mL distilled water.) [*NOTE: Some teachers prefer to use a sucrose solution* ($C_{12}H_{22}O_{11}$ *labeled* "$C_{12}O_{11}H_{22}$") *instead of acetic acid. This solution will not conduct electricity, nor will it give an acid or base indication.*]

Phosphoric acid, H_3PO_4, 500 mL (Add 5-10 mL concentrated H_3PO_4 to 500 mL distilled water.)

Sulfuric acid, H_2SO_4, 500 mL (Add 5-10 mL of concentrated H_2SO_4 to 500 mL of water.)

Advance Preparation

Prepare the solutions ahead of time. *NOTE: Specific concentrations of solutions are not important. You will want to be certain that all solutions, except sucrose, if used, conduct electricity and show positive litmus tests. Distilled water should be used. Since acetic acid and phosphoric acid are weak acids, they may show less conductivity than the others, indicating a transition from strong base to strong acid.*

Doing Chemistry DMEX C26 (Videodisc side C, Frame 25883 and Teacher's Manual p. C26-62) provides details for a microscale electrical conductivity demonstration. However, a simpler microscale electrical conductivity meter can be constructed from a battery and a low voltage light bulb. Either type is safer than using 120 volt electricity as the electrical circuit.

Materials

Wooden stick 18 cm long (paint stirrer cut in half) or Empty plastic film container with two holes drilled in the cap for the LED leads (a 9-V battery fits nicely in this container)
2 Pieces of 14- or 16-gauge Copper wire, each 8 cm long
Electrical tape
9V Battery
9V Battery Holder (Radio Shack #270-326A)
6-32 x 1/2" Screw with nut
9V Battery Clip (Radio Shack #270-325)
Light-emitting diode (Radio Shack #276-041)
330 ohm Resistor (Radio Shack #271-017)
Solder
Soldering iron
Drill with 1/8" and 1/4" bits
Wire stripper or pocket knife

Procedure

1. Drill a 1/8" hole in the wooden stick 6 cm from one end.
2. Drill a 1/4" hole in the wooden stick 6 cm from the other end.
3. Attach the battery holder to the wooden stick with the 6-32 screw and nut. (Use the 1/8" hole).
4. Solder one of the terminals of the light-emitting diode to one of the leads of the battery clip *[NOTE: The diode must be connected with the correct polarity. The polarity of the diode is usually indicated on the back of the package. The cathode (negative terminal) of the diode must be connected to the negative terminal of the battery clip (usually the black wire) or the anode (positive terminal) of the diode must be connected to the positive terminal of the battery clip (usually the red wire).]*
5. Solder one of the resistor terminals to the other lead of the battery clip.
6. Solder one of the copper wires to the other resistor terminal (if the copper wire is insulated, strip about 2 cm of insulation off of each end).
7. Solder the other copper wire to the remaining terminal of the diode.
8. Test the apparatus by attaching the battery clip to the battery and touching the two copper wires together.
9. Insert the battery in the battery holder and the light-emitting diode in the 1/4" hole.
10. Tape the copper wires to the wooden stick so that the wires do not touch each other.
11. A piezoelectric buzzer (Radio Shack) soldered in parallel to the diode will give a "sound and light" show as well.

Pre-Laboratory Discussion

Location of solutions and litmus paper should be stated. You should review the safety guidelines with students given in the Student Instructions.

Teacher-Student Interaction

1. Make sure students use clean beakers and conductivity apparatus electrodes are rinsed after each test.
2. Observe the results obtained by students. Discuss the activity with students, but do not reveal results of the various tests at this time.

Anticipated Student Results

Solution	Litmus test	Conductivity test
NaOH	Pink to blue	Conducts
MgO_2H_2	Pink to blue	Conducts
$C_2O_2H_4$	Blue to pink	Conducts
PO_4H_3	Blue to pink	Conducts
SO_4H_2	Blue to pink	Conducts

Answers to Data Analysis and Concept Development

Solution	Ions present
NaOH	Na^+, OH^-
MgO_2H_2	Mg^{2+}, $2OH^-$
$C_2O_2H_4$	H^+, $C_2H_3O_2^-$
PO_4H_3	$3H^+$, PO_4^{3-}
SO_4H_2	$2H^+$, SO_4^{2-}

NOTE: The ions may not be exactly as shown. They are written for the sake of simplicity and to emphasize the presence of H^+ *and* OH^-. *Students may have difficulty with acetic acid. Other oxy-hydrogen compounds of carbon could be used, some of which are nonconductors. Consider sugars (e.g., sucrose,* $C_{12}O_{11}H_{22}$, *a nonconductor) or carbonic acid (*CO_3H_2, *a weak conductor) as substitutes.*

Answers to Implications and Applications

1. Compounds form water solutions that are basic and ionic on the left, neutral and non-ionic in the middle, and acidic and ionic on the right side of the period.
2. Elements on the left form OH^- ions and those on the right form H^+ ions.
3. ClO_4H should conduct electricity and test as an acid (blue litmus turns pink). It should dissociate forming H^+ and ClO_4^-.

Post-Laboratory Discussion

1. Initial discussion can center around laboratory results and students' answers to the questions.
2. After discussion is completed, you may wish to reveal identities for the substances in solution, along with their usual formulas.
3. Some students may ask why aluminum was omitted from the activity. Aluminum hydroxide [$Al(OH)_3$] is insoluble in water. It will, however, dissolve in acid or base solution. This makes the compound amphoteric. This concept could be demonstrated. (Refer to the *Chemical Facts* in the *Student Version*.)

Extensions

1. *Demonstration 1* in this module.
2. *Activity 2* in the *Halogens* module involves the comparative reactivity of these elements.
3. *Group Activities* included in this module of study.

Assessing Laboratory Learning

Since this activity provides a single periodic trend, it may be best to include it with other learning assessments, rather than testing over this activity alone. Questions of the type included in the *Implications and Applications* section of this activity would be good to use as an evaluative instrument.

Suggestions for Other Laboratory Activities

Relative Reactivity of Halogens and the Halides (*Activity 2* in the *Halogen* module).

DEMONSTRATIONS

CAUTION: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Relative Reactions of Alkali Metals with Water

Introduction

This demonstration should be used along with *Activity 1*. As an alternative, consider using a videotape (see *Media*). Because of the danger of storing and handling potassium, it is recommended that a video demonstration of potassium reactivity be used along with this demonstration to extend the periodic trend.

The purpose of this demonstration-discussion activity is to show relative reactivities of alkali metals by reacting two of them (Li and Na) with water. [Use the videodisc on the Periodic Table (see *Reference*) to illustrate the reaction of potassium with water.] These activities should be performed as a demonstration rather than as a student activity because of the high reaction rates, the somewhat variable way in which the reactions occur, and the dangers involved in handling metals improperly.

This demonstration is effectively done as a step-by-step process in which students write observations, equations, and answer questions while the demonstration is being performed. The process will require one class period.

Safety

These metals should be kept in separate beakers, and the metals should be covered with dry kerosene. The beakers should be covered with watch glasses or other appropriate covers. *This demonstration should be performed before the class only after you have practiced it carefully. Observe all safety considerations including wearing safety goggles.*

Materials

Lithium metal, Li, 2 pea-sized pieces
Sodium metal, Na, 3 pea-sized pieces
Pair of forceps
Pair of crucible tongs
Test-tube, 18- x 150-mm
Square of aluminum foil, 5- x 5-cm
2 Beakers, 400-mL
2 Watch glasses (or other appropriate covers for the reaction beakers.)
Litmus paper
5% Phenolphthalein (in 95% ethanol)
Paper towels

Procedure

For convenience, a list of questions is submitted. This list may be copied on a handout so that students can answer the questions after viewing the videodisc and as the demonstration progresses.

1. Using forceps, remove one piece of sodium from the kerosene. Dry the metal by pressing it in folds of paper towel. After testing the water with litmus paper, drop the piece of sodium in the water and immediately cover beaker with a watch glass. After reaction is finished, remove the watch glass and test solution with litmus paper and with phenolphthalein.

   NOTE: Students often want to see the reaction repeated. If that is done, fresh water should be obtained for the second reaction.

2. Put another piece of dried sodium on a 5- x 5-cm square of aluminum foil, and fold the foil around it so that there is a small opening to allow gas to escape. Holding the foil with crucible tongs, submerge the foil in another beaker of water, and invert the 18- x 150-mm test-tube filled with water so that some of the gas evolved can be collected in the test-tube. The test-tube can be left inverted in the beaker until it is convenient to perform the flammability test. After students have answered Question 1, have them share their observations.

3. Test the gas for flammability. After students agree that hydrogen gas is produced, have them answer Question 2.

4. Point out the location of lithium on the Periodic Table. Ask students to predict how lithium's reaction rate will compare to that of sodium, and then perform the demonstration with lithium. (*NOTE: The handling of Li and the reaction system are the same as for Na.*) It is not essential that the class have consensus on lithium's rate before the reaction, but differing hypotheses should be supported with logical reasoning. After the first reaction with lithium, it may be necessary to use another piece of lithium in the same container in order to get a clear litmus test. Have students write a balanced equation for this reaction (Question 3).

5. Instruct students to continue answering Questions 4 and 5. These questions can be discussed until general consensus is reached.

6. Oral discussion of Question 6 should precede viewing the video of the reaction of potassium with water. Once again, consensus is not as important as reasoned hypothesis.
7. Watch the video demonstration of the reaction of potassium with water. Have students write a balanced equation for the reaction (Question 7).
8. If time permits, ask students to answer Question 8, and discuss their answers. The question may be given as a homework assignment.

Questions

1. What evidence indicates that the combination of sodium and water produced a chemical change? *[Gas produced, litmus changed to blue, phenolphthalein turned red, heat]*
2. After your teacher collects and tests the gas produced, write a balanced equation for the reaction of sodium with water.

   $$[2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)]$$
3. Write a balanced equation for the reaction of lithium with water.

   $$[2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)]$$
4. How does the reactivity of lithium with water compare to that of sodium with water. *[The reaction of sodium is greater, indicating a greater reactivity for sodium.]*
5. Offer an explanation for the differing reactivities of lithium and sodium in terms of electron arrangements of the two elements. *[The electron configurations of Li and Na are: Li 2–1 and Na 2-8-1. Sodium loses its valence electron more readily than does lithium. The greater distance of sodium's valence electron from its nucleus and the additional energy level of inner shell electrons in sodium combine to make the removal of sodium's valence electron easier than the removal of lithium's valence electron.]*
6. Predict potassium reactivity in comparison to that of lithium and sodium. *[After brief discussion, the class should predict that potassium will be more reactive than the others tested.]*
7. Write a balanced equation for the reaction of potassium with water.

   $$[2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)]$$
8. Based on your experience with Li, Na, and K, how would you expect the reactivity of rubidium and cesium to compare with each other and with the three metals you have seen tested? *[Assuming the reactivity trend continues, for reasons argued earlier relating to electronic structure, cesium should be more reactive than rubidium, and both of these should be more reactive than the first three metals.]*

Reference

The Periodic Table Videodisc: *Reaction of the Elements*. Journal of Chemical Education: Software (Department of Chemistry, University of Wisconsin, 1101 University Avenue, Madison, WI 53706.). This is a 30-min videodisc with still shots and action shots of each element and its common reactions. It can be used with various computer programs. Several issues of this software journal have computerized databases for the Periodic Table for various computer types. For IBM: Issue IB, 1; KC? Discoverer: Exploring the Periodic Table. For Macintosh: Issue IC,1 The Periodic Table Stack. For Apple: Issue IA, 1 The One-Computer Classroom (Periodic Table Works).

Relative Reactivity of Halogens and the Halides (*Activity 2* in the *Halogen* module) can be performed as a demonstration if it is not used as a laboratory activity.

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. What trends (*e.g.*, in atomic size and ionization potential) are shown among elements as one moves across a period of the Periodic Table? *(Because the outer electrons are held more strongly as one moves across a period of elements, atomic size would decrease and ionization potential would increase.)*
2. What trends (*e.g.*, atomic size and ionization potential) are shown among elements as one moves down a column of the Periodic Table? *(Because outer electrons are held more weakly as one moves down a column of the Periodic Table, atomic size would increase and ionization potential would decrease.)*
3. Account for the similarity of properties among families of elements in the Periodic Table. *(Identical numbers of outer electrons in the electron structure among a family of elements account for similarities in properties.)*

Counterintuitive Examples and Discrepant Events

1. Examine "non-periodic" trends in ionization energies of Period 3 elements. There are "blips" in the trend that can be explained in terms of electron structure (changes in s to p electrons, changes in p^3 to p^4 electrons). *[One might expect that ionization energies would increase steadily across a period, because of the greater number of protons with each successive element and the low electron shielding effect from adding electrons to the same energy level. With two exceptions in each period this pattern exists. It can be noted that removing an electron from a full s-sublevel provides one of the "blips," and removing an electron from a half-filled p-sublevel provides the other "blip." It can be stated, then, that filled and half-filled sublevels appear to have a higher stability, with respect to electron removal, than do other conditions.]*
2. Discuss hydrogen as an example of an alkali metal. *[Since hydrogen is a gas that exists as diatomic molecules in the elemental state, one might consider it odd that many periodic tables list hydrogen above the alkali metals. There are similarities to the alkali metals: hydrogen has one valence electron, hydrogen has positive character in many compounds, and hydrogen sometimes forms positive ions in solutions. Hydrogen's comparatively high ionization energy and electronegativity cause it to behave in some cases, as a halogen, tending to form negative ions (hydrides) in compounds. Furthermore, two of the halogens are gases at room temperature, as is hydrogen. Some forms of the Periodic Table list hydrogen with the alkali metals and halogen families.]*
3. Discuss the appearance of the Periodic Table if atomic weight rather than atomic number had been used to arrange the chart. How does one explain the Co-Ni, Ar-K, and Te-I reversals? *[In these cases of atomic weight "discrepancy" the most commonly occurring stable isotopes of Ni, K, and I are lighter than the most commonly occurring stable isotopes of Co, Ar, and Te. These reversals could be related to the observation that, of the naturally occurring stable isotopes, the most stable ones have an even number of protons and an even number of neutrons—"even, even." Next, in stability, would be "odd, even," and "even, odd." There are very few naturally occurring stable isotopes that have both an odd number of protons and an odd number of neutrons.]*

4. Examine the periodic and non-periodic behavior of some transition metal families. *[Transition metal behavior could be assigned as a project for individuals or small groups of students who are interested in pursuing possible periodic patterns beyond looking at representative elements. Graphs of 1st and 2nd ionization energies for Periods 4 and 5 show some interesting similarities, which could produce useful discussion and patterns relating to electron arrangements. Compiling common oxidation states for these elements and examining them with regard to periods and families could be constructive.]*

Metaphors and Analogies

The following activity (Xeno simulation) is taken from Dorin, Demmin and Gabel. *Concept Mastery, Chemistry: The Study of Matter*, Chapter 14, p. 385. Boston: Prentice-Hall, and from the Teacher's Guide, p. TG-212.

Have students imagine they are scientists on another world, where there is a set of elements different from Earth's. The name of the planet is Xeno. The students' job is to perform tests on the elements of Xeno, and to arrange the elements in a Periodic Table with families and periods, similar to the Periodic Table used on earth. They obtain the information in stages.

Directions

Stage 1 Testing: Collect data on the physical properties of the elements. These data are given in the following table. Devise a Periodic Table based on these data. Justify your groupings.

Physical Properties of Elements on Xeno*

Element	Color	Hardness	Melting Point (°C)
A	Turquoise	Soft	1050
B	Silvery, black	Hard	–300
C	Yellow	Soft	1000
D	Gray	Hard	400
E	Pink	Soft	1200
F	Silvery, black	Hard	–100
G	Silvery, black	Hard	–200
H	Black	Hard	300
I	Aqua	Soft	900
J	Brown	Soft	1000

*Room temperature on Xeno is –320 °C.

Stage 2 Testing: Collect data on the chemical properties of the elements. Using the data listed below, modify your original periodic table. Justify your new arrangement.

Chemical Properties of Elements on Xeno

Element	Reacts with water	Reacts with acid	Reacts with oxygen	No reaction
A			X	
B	X	X	X	
C				X
D		X	X	
E				X
F	X	X	X	
G	X	X	X	
H	X	X	X	
I			X	
J		X	X	

Stage 3: Determine relative atomic masses for elements on Xeno. Using the data listed below, modify your periodic table. Justify your new arrangement.

Relative Atomic Masses of Elements on Xeno

Element	Relative atomic mass	Element	Relative atomic mass
A	5	F	15
B	3	G	9
C	1	H	14
D	7	I	1
E	10	J	6

Teaching Notes

Stage 1: The first arrangement might look like the table that follows. Notice that melting point increases across a period and down a group, whereas hardness decreases. (At this stage, it is equally valid to arrange the elements so that melting point decreases across a period or down a group.)

The elements on the left are all silvery and black; those in the middle, darkly colored; and those on the right, brightly colored. (With the given information, element J could also be put in the right group, either above or below C.)

B	–333
G	–200
F	–100

H	300
D	400
J	1000

I	900
C	1000
A	1050
E	1200

Stage 2: The table reflects similarities in chemical properties. Since B, G, and F show the same chemical properties, the left group remains intact. The middle group is split in two with H on the left because its chemical properties are identical to those of the elements in the left group. D and J do not react with water, so they are closer to the elements in the right group.

B	H	D	I	C
G		J	A	E
F				

Stage 3

C	I				B
	A	J	D		G
E				H	F

a. Notice that the brightly colored elements are now on the left side of the table.

b. Xeno's unreactive elements (C and E) are on the left side of the table, whereas earth's noble gases are on the right. Xeno's hard, silvery elements would be gases at 25 °C (earth's room temperature) and are on the right, whereas earth's metals are solid at 25 °C and are on the left side of the table.

c. This table now has 3 periods and 6 families (groups).

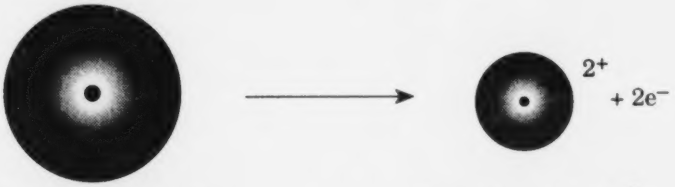
d. This table has places for 18 elements, 10 of which are known. The properties of a missing element should follow the trends in properties within its period and its group.

e. Although it is easy to record the physical properties given in Stage 1, these properties by themselves do not give enough information to build a periodic table. The chemical properties from Stage 2 are very useful, but the best information is the relative atomic masses from Stage 3.

Notice the similarties between this process of building your table for elements on Xeno and the historical development of the Periodic Table on Earth.

Pictures in the Mind

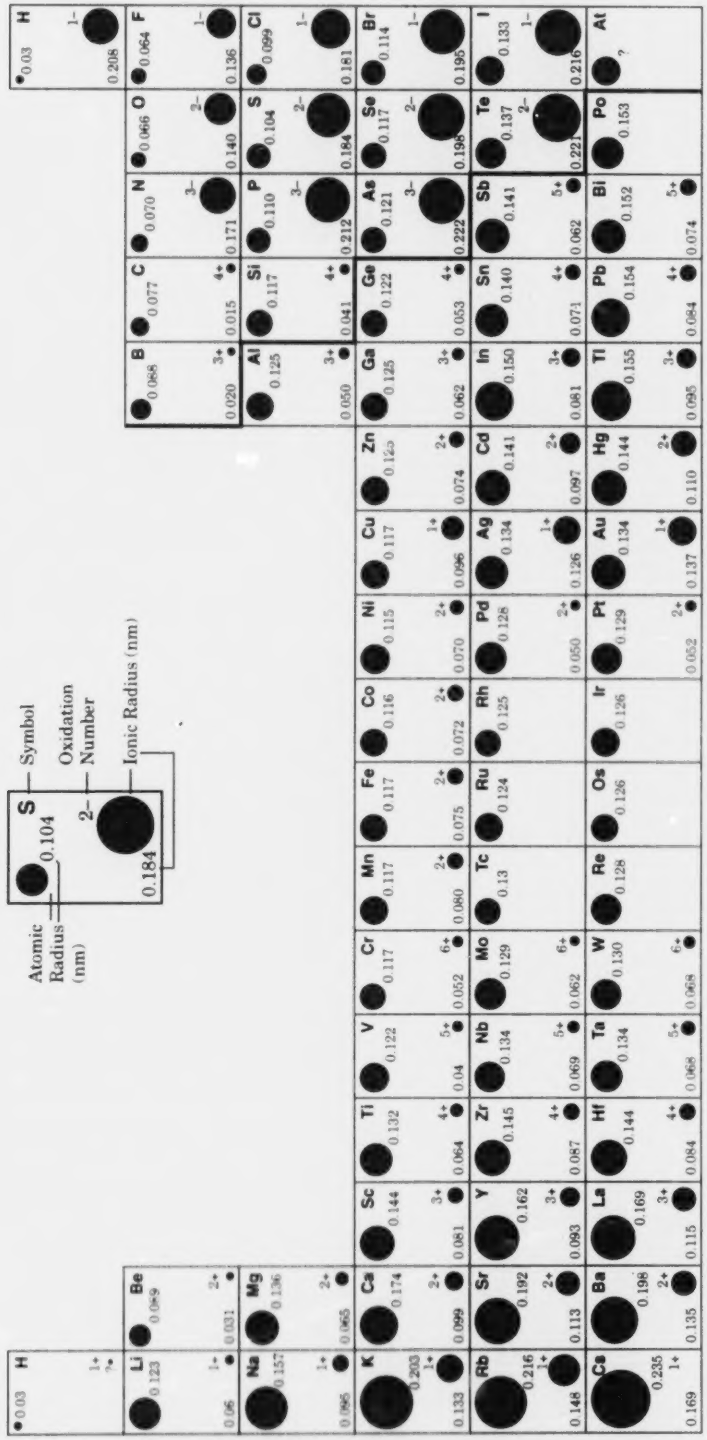
1. Have students draw circles to represent relative sizes of atoms in selected families and periods of the Periodic Table.
2. Have students draw circles to represent relative sizes of atoms and ions of selected elements. [See *Appendix* from Dorin, Gabel, and Dennmin. (1989). *Chemistry: The study of matter*. Prentice-Hall, pp. 376-377, 379.]



Magnesium atom, Mg
(radius 0.136 nm)

Magnesium ion, Mg^{2+}
(radius 0.065 nm)

Figure 2. Relative sizes of Mg and Mg^{2+}.


S
0.104
2–
0.184
Symbol
Oxidation Number
Ionic Radius (nm)
Atomic Radius (nm)
H 0.03 1+ ?
Li 0.123 1+ 0.06
Be 0.089 2+ 0.031
Na 0.157 1+ 0.095
Mg 0.136 2+ 0.065
K 0.203 1+ 0.133
Ca 0.174 2+ 0.099
Sc 0.144 3+ 0.081
Ti 0.132 4+ 0.064
V 0.122 5+ 0.04
Cr 0.117 6+ 0.052
Mn 0.117 2+ 0.080
Fe 0.117 2+ 0.075
Co 0.116 2+ 0.072
Ni 0.115 2+ 0.070
Cu 0.117 1+ 0.096
Zn 0.125 2+ 0.074
Rb 0.216 1+ 0.148
Sr 0.192 2+ 0.113
Y 0.162 3+ 0.093
Zr 0.145 4+ 0.087
Nb 0.134 5+ 0.069
Mo 0.129 6+ 0.062
Tc 0.13
Ru 0.124
Rh 0.125
Pd 0.128 2+ 0.050
Ag 0.134 1+ 0.126
Cd 0.141 2+ 0.097
Cs 0.235 1+ 0.169
Ba 0.198 2+ 0.135
La 0.169 3+ 0.115
Hf 0.144 4+ 0.084
Ta 0.134 5+ 0.068
W 0.130 6+ 0.068
Re 0.128
Os 0.126
Ir 0.126
Pt 0.129 2+ 0.052
Au 0.134 1+ 0.137
Hg 0.144 2+ 0.110
B 0.088 3+ 0.020
C 0.077 4+ 0.015
N 0.070 3– 0.171
O 0.066 2– 0.140
F 0.064 1– 0.136
Al 0.125 3+ 0.050
Si 0.117 4+ 0.041
P 0.110 3– 0.212
S 0.104 2– 0.184
Cl 0.099 1– 0.181
Ga 0.125 3+ 0.062
Ge 0.122 4+ 0.053
As 0.121 3– 0.222
Se 0.117 2– 0.198
Br 0.114 1– 0.195
In 0.150 3+ 0.081
Sn 0.140 4+ 0.071
Sb 0.141 5+ 0.062
Te 0.137 2– 0.221
I 0.133 1– 0.216
Tl 0.155 3+ 0.095
Pb 0.154 4+ 0.084
Bi 0.152 5+ 0.074
Po 0.153
At ?
H 0.03 1– 0.208

Adapted from Dorin, H., Demmin, P., and Gabel, D. (1989). *Chemistry: The study of matter.* Prentice Hall, Inc.

Figure 3. Covalent Atomic radii and ionic radii.

Other Activities

1. What country has been named for an element? (*HINT*: It must be an "old" element and a "new" country. Why?) *[Although many elements are named after a country, only one country, Argentina, is named after an element silver. The name for silver is from the Latin argentum.]*
2. Periodic Chart Puzzles based on the game Hang-Man or Wheel of Fortune. Examples that follow are taken from Stafford, D. G. *et al.* (1987). *Top chem: The Oklahoma project: Chemistry*. Oklahoma City: Oklahoma State Department of Education.

The primary purpose of the Periodic Chart Puzzles is to provide an opportunity to reason with concepts, patterns, and models. This is the essence of scientific thinking. The patterns that you will be using are the trends of property, size, periods (rows) and groups (columns). These patterns and model constitute an important part of the structure of chemistry. If you can solve these puzzles, you have a good grasp of some of the important basic building blocks of chemistry.

The concepts and patterns could relate to many concepts discussed in this module and in others: *s*, *p*, *d*, *f* elements, electron configuration, orbital diagram, ionization potential, electronegativity, atomic size, isoelectronic series, relative atomic weight, orbital diagram, metal, nonmetal, period, group, family, isotope, valence electron, ionic size, quantum numbers, atomic number, unpaired electrons, empirical formula and metalloid.

The puzzles should be fun to do! As each puzzle is worked try to see the concept or pattern clearly in your mind. It is in this way that you will gain most from the activity. Get help only if you cannot reason out the answer for yourself.

An activity that can be even more fun is creating puzzles of your own. Make a puzzle and exchange it with one made by someone else. This can be a very educational activity. The creative aspect of logical thinking is very important in science.

In many cases, both letters in the symbol are used. When that is the case, the numbers in the adjacent blocks are the same. For example, if the word is *Site*, then the four blocks would be:

1	1	2	2

For the elemental symbols

Si Te

"Thing" by Don Stafford

Clues:

1	2	3	4	5	5	6	6	7	8	8	9
___	___	___	___	______		______		___	______		___

1. Boron family, Period 3, first letter
2. Atomic Number 5
3. Oxygen family, Period 3
4. Oxygen family, Period 2
5. Atomic Number 71
6. Oxygen family, Period 5
7. Atomic Number 40, first letter
8. Atomic Number 68
9. Atomic Number 8

Answer: ABSOLUTE ZERO

"Thing" by Connie Sloan

1 1 2 3 4 5   6 7 8   9 10 10 11

Clues:

1. 84 protons
2. 53 protons
3. 16 protons
4. Atomic Number 8
5. Z = 7
6. Z = 53
7. Z = 23
8. Z = 39
9. Atomic mass = 183.85
10. Atomic mass = 26.98
11. Atomic mass = 39.10

Answer: POISON IVY WALK

"Phrase" by Donna Deen

1 2   3 4 4   5 6 7   8 8 9 10 10

6 9   11 11 12 12 13   14 6   15 15

9 16 17 17 3   6 9   18 6 3   19 19

20 20 13

Answer: BE THE JOB LARGE OR SMALL
DO IT RIGHT OR NOT AT ALL

Clues:

1. Boron family, Period 2
2. Atomic Number 4, second letter
3. Atomic mass 232, first letter
4. Lightest noble gas.
5. Tenth letter in the alphabet.
6. Atomic mass of 15.999
7. Atomic Number 5
8. Element for which Lanthanide series is named.
9. *spdf* configuration of $1s^2 2s^2 2p^6 3s^2 3p^6$ second letter
10. Element that contains 32 electrons.
11. At. No. = 62
12. Period 3, oxidation number of +3
13. Atomic Number 3, first letter
14. Atomic Number 101, second letter
15. Atomic Number 22, reversed
16. A solid halogen mistakenly thought of as a liquid.
17. Element known as quicksilver, reversed.
18. Heaviest noble gas, second letter.
19. Halogen family element with lowest electronegativity.
20. Boron family, $3p^1$ element

"PERIODICALLY PUZZLING"

[Nelson, B. J. (1985, December). ChemMatters, p. 16.]

Here are some questions about the elements, their properties, and periodicity. In each case, the correct answer is the name of an element (one of the two given). Write the answer in the appropriate space at right, placing the first letter in the first blank. There are more than enough blanks to fit either element name, so the number of letters in the name cannot be used as a clue. When the puzzle is complete, the Periodic Law will be displayed in the shaded column.

1. More outer electrons, phosphorus or silicon?
2. Liquid at room temperature, cobalt or mercury?
3. Produces a yellow flame test, barium or sodium?
4. Can be cut with a knife, dysprosium, or chromium?
5. Part of the lanthanide series, cerium or thorium?
6. Larger atoms, carbon or fluorine?
7. Very high ionization energy, krypton, or potassium?
8. Less dense, strontium or lithium?
9. A 1^+ or 1^- charge, hydrogen or rubidium?
10. Named after the theorizer of relativity, curium or einsteinium?
11. Configuration ends in $3p^1$, aluminum or sulfur?
12. Gives a yellow-green flame test, potassium, or barium?
13. Named after an Eastern Hemisphere continent, americium, or europium?
14. Has a dot diagram of X·, calcium or potassium?
15. Forms a hexafluoride, argon or xenon?
16. A metal used in jewelry, sodium, or rhodium?
17. Isolated by Madame Curie, radium or rhenium?
18. Its crystals sublime as violet gas, iodine or bromine?
19. Brass mimics this metal, copper, or gold?
20. One of the two most dense metals, osmium, or magnesium?
21. Used to disinfect water, boron or chlorine?
22. Compounds of this are used in toothpaste, lithium or fluorine?
23. Named after the inventor of dynamite, lawrencium or nobelium?
24. Has a higher ionization energy, antimony, or bismuth?
25. Has the higher oxidation state, scandium, or calcium?
26. Is smaller, titanium or zirconium?
27. Reacts with H_2O to produce H_2, calcium, or copper?
28. Atomic number 5, boron or nitrogen?
29. Used in the breeder reactor, uranium, or plutonium?
30. Used to make computer chips, krypton or silicon?
31. Aqueous solutions are blue, copper or nickel?
32. Low melting point, francium or platinum?
33. A poison; carbon or arsenic?
34. Forms a 2^+ ion, strontium or boron?
35. Configuration ends in $6p^6$, radium or radon?
36. Known for heavy metal poisoning, cadmium or magnesium?
37. Becomes rust, iron or nickel?
38. Deep blue compounds, cobalt or zinc?
39. Aqueous solutions are green, nickel or sodium?
40. A yellow element, selenium or sulfur?
41. Atomic weight 165, terbium or holmium?
42. Alkali metal, period 5, rubidium or strontium?
43. Used as light bulb filaments, platinum, or tungsten?
44. Reactor fuel, uranium or nobelium?

1 _ _ _ _ _ _ _ _ _ _ _
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3 _ _ _ _ _ _ _ _ _ _ _ _
4 _ _ _ _ _ _ _ _ _ _ _ _ _ _
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37 _ _ _ _ _ _ _ _ _ _ _
38 _ _ _ _ _ _ _ _ _ _ _

39 _ _ _ _ _ _ _ _ _ _ _
40 _ _ _ _ _ _ _ _ _ _ _ _
41 _ _ _ _ _ _ _ _ _ _ _ _ _ _
42 _ _ _ _ _ _ _ _ _ _ _ _ _
43 _ _ _ _ _ _ _ _ _ _ _ _ _ _ _ _ _
44 _ _ _ _ _ _ _ _ _ _ _ _

Answer to "Periodically Puzzling": Properties are periodic functions of atomic number.

1 PHOSPHORUS
2 MERCURY
3 SODIUM
4 DYSPROSIUM
5 CERIUM
6 CARBON
7 KRYPTON
8 LITHIUM
9 HYDROGEN
10 EINSTEINIUM

11 ALUMINUM
12 BARIUM
13 EUROPIUM

14 POTASSIUM
15 XENON
16 RHODIUM
17 RADIUM
18 IODINE
19 GOLD
20 OSMIUM
21 CHLORINE

22 FLUORINE
23 NOBELIUM
24 ANTIMONY
25 SCANDIUM
26 TITANIUM
27 CALCIUM
28 BORON
29 PLUTONIUM
30 SILICON

31 COPPER
32 FRANCIUM

33 ARSENIC
34 STRONTIUM
35 RADON
36 CADMIUM
37 IRON
38 COBALT

39 NICKEL
40 SULFUR
41 HOLMIUM
42 RUBIDIUM
43 TUNGSTEN
44 URANIUM

"IT'S ELEMENTAL"

[Lowery, J. O. (1982, December). ChemMatters, p. 16.]

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M E R C U R Y B A R I U M U I P N E N I T A T S A
M U I C L A C N I O B I U M U R E M U I H T I L N
U R I A R D O E L H E L I U M O P C A D M I U M T
I O U M A R G O N E S U O U Y T T E R Z T M E N I
C D M I O D I N E T S G N U T A U N A I I U M E M
I A L B X R O T H O R I U M M C N I D N N S S U O
R N A O Y E H T H O T U O N A T I C U C O I C T N
E S B L G R N C B A L M S O D I U M U I L L A G Y
M U O T E B I O L R L U B D O N M E M V A I R O N
A R S E N I C P N G L L N A G I A N E M N C B X B
C O H E N U K P L F B E I R M U G R U S T O O Y E
T H M O M M E E U C E S I U M M N N L K H N N G R
S P E M D U L R A O R I N U M E E S M R A D I U M
M S N A Y I I U P N Y U I B N D S U L Y N Y T M R
U O D N T N U M E O L M O I B R I I E P U S R U U
I H E G T A L M S R L I R Y H C U T C T M P O I B
N P L A R M R N I O I O L E N C M T A O R B G N I
I L E N I R S N H B U O N E O B U U H N B L E A D
E U V E U E E E A L M I R I R O I R I E T A N T I
T T I S M G L P F E U W I O U R S S I L N A L I U
S O U E O D E T N M A N M B O M S O M U E I L T M
N N M R A D N U I L D I N H U R A N I U M K U U M
I N D I U M I M U I N O T U L P T B R I T O R M M
E Y T T E R U U M E M U I L E B O N L O W H I E N
H C E R I U M M C H L O R I N E P F E R M I U M B
```

Find the full name of the elements identified below by their chemical symbols. Names can be backwards, in rows, in columns and diagonally. Circle each as you find it.

Ag Al Am Ar As At Au B(2) Ba Be Bi

Bk Br C Ca Cd Ce Cl Cm Co Cr Cs

Cu Er Es F Fe Fm Ga Ge H He Hf

Hg Ho I In K Kr La Li Lr Md Mg

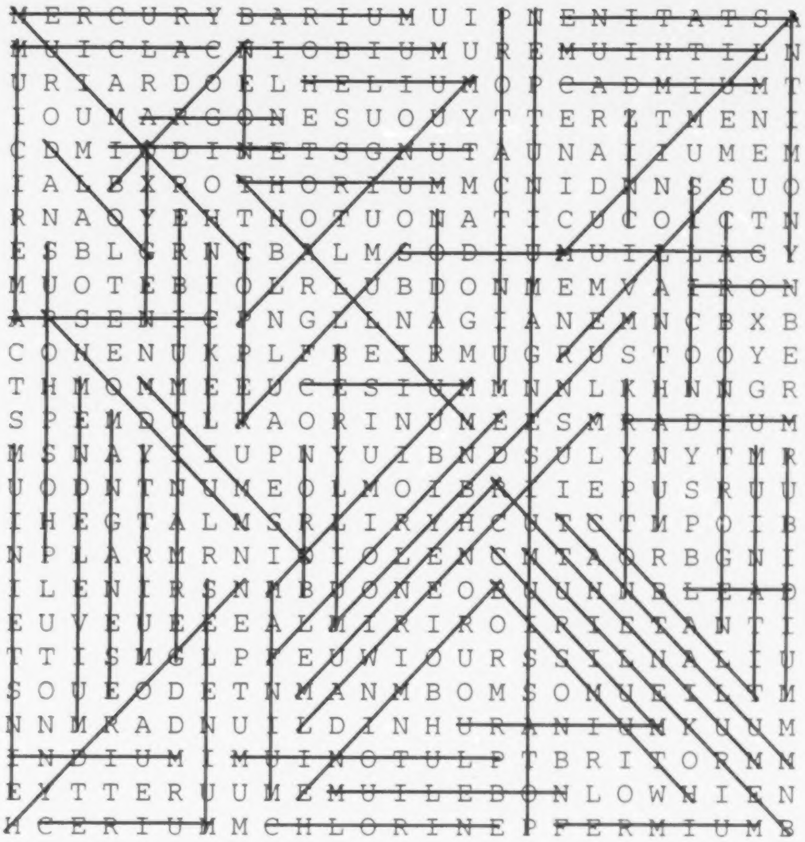
Mn Mo N Na Nb Ne Ni No Np O Os

P Pa Pb Po Pt Pu Ra Rb Re Rh Rn

Ru S Sb Se Si Sn Ta Th Ti Tl U

W Xe Y Zn

Answer key to "It's Elemental":



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M E R C U R Y B A R I U M U I P N E N I T A T S A
M U I C L A C N I O B I U M U R E M U I H T I L N
U R I A R D O E L H E L I U M O P C A D M I U M T
I O U M A R G O N E S U O U Y T T E R Z T M E N I
C D M I O D I N E T S G N U T A U N A I I U M E M
I A L B X R O T H O R I U M M C N I D N N S S U O
R N A O Y E H T N O T U O N A T I C U C O I C T N
E S B L G R N C B X L M S O D I U M U I L L A G Y
M U O T E B I O L R E U B D O N M E M V A I R O N
A R S E N I C P N G L E N A G I A N E M N C B X B
C O H E N U K P L F B E I R M U G R U S T O O Y E
T H M O N M E E U C E S I U M M N N L K H N N G R
S P E M D U L K A O R I N U N E E S M R A D I U M
M S N A Y I I U P N Y U I B N D S U L Y N Y T M R
U O D N T N U M E O L M O I B R I I E P U S R U U
I H E G T A L M S R L I R Y H C U T C T M P O I B
N P L A R M R N I D I O L E N C M T A O R B G N I
I L E N I R S N M B U O N E O B U U H N B L E A D
E U V E U E E E A L M I R I R O L R I E T A N T I
T T I S M G L P F E U W I O U R S S I L N A L I U
S O U E O D E T N M A N M B O M S O M U E I L T M
N N M R A D N U I L D I N H U R A N I U M K U U M
I N D I U M I M U I N O T U L P T B R I T O R N N
E Y T T E R U U M E M U I L E B O N L O W N I E N
H C E R I U M M C H L O R I N E P F E R M I U M B
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"FAMILY RESEMBLANCE"

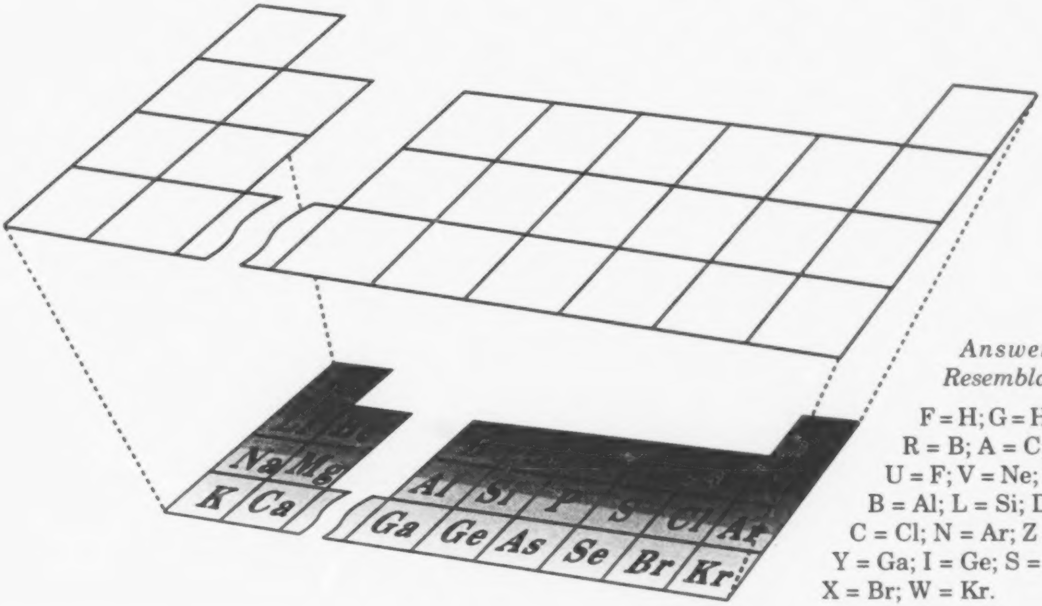
[Wolfgram, D. (1987, December). *ChemMatters,* p. 16.]

Can you place these elements in their proper positions on the Periodic Table without knowing their identity? Instead of the usual chemical symbol, each element is represented by a code letter, and some clues to the element's identity are listed below. In addition, the elements fall into these families or groups:

JKQ, WNVG, IAL, EPFZ, YBR, ODS, CXU, THM

Place each code letter in its proper space in the table below, then indicate the actual chemical symbol that corresponds to each letter.

W is an inert gas.
P is an alkali metal.
X is a halogen.
D is an element that needs three electrons to become stable.
K is an atom that readily yields a 2+ ion.
M is an element found in the group to the right of S.
B would probably form a 3+ ion.
A has six protons in its nucleus.
R has an atomic number one less than A.
O has an atomic number one more than A.
S has 18 more protons in its nucleus than has D.
F has only one proton in each atom.
Y has atoms that are larger than the atoms of B.
I has the largest atomic number in its group.
T is found in the fourth period (or row).
C has atoms that are larger than those of U, but smaller than those of X.
V has one more proton in its nucleus than U.
H has the smallest atomic number in its group.
G has two protons in each atom.
N is a third period element.
J is a second period element.
Q has an atomic mass greater than that of J and less than that of K.
E has an atomic number less than that of Z but greater than that of P.
Z has more protons than other elements in its group.
U is found in the second period.



Answers to Family Resemblance:

F = H; G = He; P = Li; J = Be;
R = B; A = C; O = N; H = O;
U = F; V = Ne; E = Na; Q = Mg;
B = Al; L = Si; D = P; M = S;
C = Cl; N = Ar; Z = K; K = Ca;
Y = Ga; I = Ge; S = As; T = Se;
X = Br; W = Kr.

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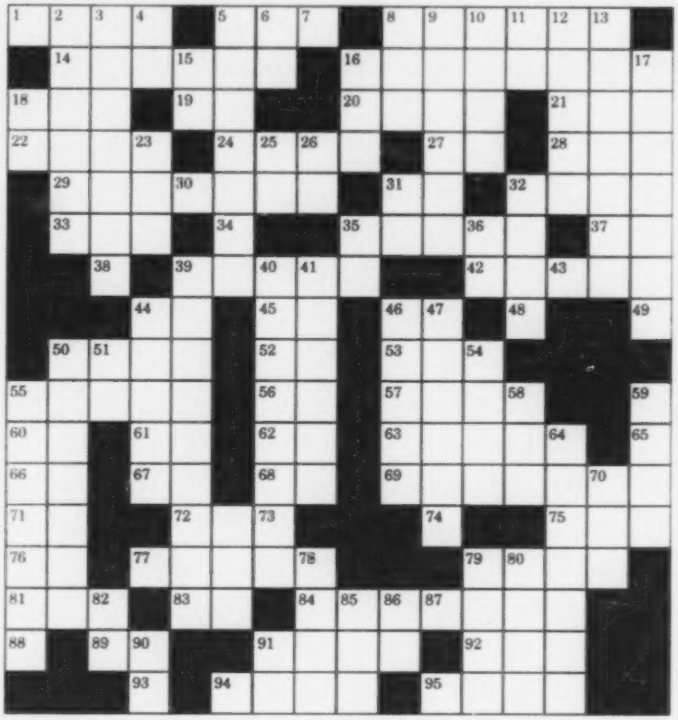
"IT'S ELEMENTARY"

[Landon, L. (1989, February). *ChemMatters*, p. 16.]



Across

1. Product of an acid and base.
5. Black mixture of hydrocarbons.
8. HIO_2 is ____ acid.
14. Element used in coins.
16. Bond angles too large or too small.
18. Massachussetts Institute of Technology, abbr.
19. 1/1000 of 15 down, var.
20. Given off by an exothermic reaction.
21. Solid phase of water.
22. Danish atomic theorist, Niels ____.
24. A substance that produces hydroxide ions in water.
27. Element; ancient name natrium.
28. 2000 lbs.
29. Element formerly named columbium; renamed after daughter of Tantalus.
31. Element; ancient name, argentum.
32. Synthetics, abbr.
33. Total.
34. Chemical symbol for element number 92.
35. See 9 down.
37. Chemical prefix denoting two.
38. Meter.
39. An organic compound containing $-NH_2$ group.
41. Noble gas used in advertising.
42. Humid.
44. Element named after brilliant indigo line in its spectrum.
45. See 41 across.
46. Eighth most abundant element in earth's crust.
48. Element once called wolfram.
49. ____ axis.
50. Smoke and fog.
52. Heaviest of the alkali metals.
53. Electroencephalogram.
55. Can be covalent or ionic.
56. Noble gas, not 31 down or 45 across.
57. Catches vacuum filtrate or animals.
59. Symbol often used to denote entropy.
60. Most abundant metal in earth's crust.
61. Institute of Technology, abbr.
62. Element forms poisonous compounds.
63. compound with $-CONH_2$ group.
65. Angstrom.
66. Portion of the electromagnetic spectrum.
67. Element named after Greek word for color.
68. Room temperature, abbr.
69. Component of many soaps and salves.
71. See 31 down.
72. Suffix meaning a carbohydrate.
74. Moles, in universal gas equation.
75. Not the beginning.
76. See 44 across.
77. See 63 across.
79. Common name for calcium oxide.
81. Teletype, abbr.
83. Element used to plate cans to prevent corrosion.
84. Name for monatomic ion formed by nitrogen.
88. Energy, abbr.
89. Element in emeralds.
91. Not warm.
92. Writing implement.
93. See 34 across.
94. High-temperature oven.
95. Possible shape of cyclohexane ring.

Down

1. Element known in ancient days as brimstone.
2. Negatively charged ions.
3. Lightest metal.
4. The only one of the first 83 elements not found in nature.
5. One of four elements named after Swedish town.
6. See 60 across.
7. Symbol commonly used for ideal gas law constant.
8. Chemical suffix.
9. Part of ROY G BIV.
10. Information gathered in laboratory experiment.
11. Two-digit binary code for Number 1.
12. One.
13. SI time units.
15. SI unit of mass.
16. Standard hydrogen electrode, abbr.
17. Physical property related by mass and volume.
18. Megabyte.
23. Read only memory.
25. Element; ancient name aurum.
26. Symbol for samarium.
30. Element makes pyrotechnics green.
31. Chemical symbol for a noble gas.
32. Crystalline form of water.
36. Electromagnetic.
39. A unit used to define bond lengths.
40. Portion of the electromagnetic spectrum.
41. German scientist known for an equation used in electrochemistry.
43. Element that is purple in the gas phase.
44. HIO_3 is ____ acid.
46. Examples: Fe, Ag.
47. Nationality of Wilhelm Roentgen.
50. Substance that dissolves a solute.
51. Element that gives the purple color to amethyst.
54. Opposite of loss.
55. Aluminum ore.
58. Partial differential overlap, abbr.
59. A starting material in the manufacture of glass.
64. Building block of compounds.
70. Chemical suffix.
73. Vice or wrong-doing.
77. See 65 across.
78. Organic compound has $-OH$ next to $C=C$.
79. Prefix for fatty.
80. Innovative thought.
82. Like 5 down, but another town.
85. Charged atom.
86. Element has green spectral line and Greek name for green twig.
87. See 7 down.
90. Element named after a continent.
91. Roman numeral 101.
94. Element makes flame violet.
95. See 30 down.

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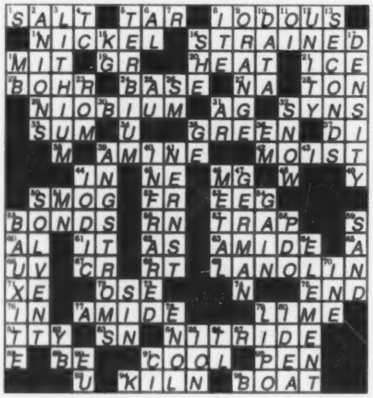
Answer to It's Elementary:



3. Have students graph relative ionization energy vs. atomic number (from data tables provided) and discuss patterns and trends. Data for this activity and suggestions for discussion follow.

   This activity has been effective when the first ionization energies are graphed and discussed prior to the graphing of second and higher ionization energies. A brief discussion of the definition of ionization energy (or ionization potential) will be helpful. (First ionization energy is the amount of energy required to remove one electron from a neutral atom that is in the gaseous state. Second ionization energy measures removal of one electron from an ion with a 1+ charge, and so forth).

   Graphs are made by plotting atomic number (Z) on the horizontal axis of the graph and relative ionization energy on the vertical axis of the graph. If students are to plot all five ionization energies on the same graph, the relative ionization energy coordinate will need to be 100 units high. Relative ionization energy is used because it allows for the use of smaller numbers. Ionization energies are all made relative to rubidium's first ionization energy so they can all be compared on the same graph. There is no need to skip ten spaces on the atomic number coordinate between 20 and 31 or between 38 and 49.

   The larger number of elements plotted on the first ionization energy graph is intended to allow students to observe the trends and patterns for more periods of elements and to see that the general trends and patterns remain the same. Of course, these data are graphed only for the representative elements.

   Suggested discussion questions follow.

First Ionization Energies

a. What patterns can you observe in the first ionization energies of elements? *[A repeating pattern of peaks and valleys defines the values of ionization energies.]*

b. How do patterns you observed relate to the placement of elements on the Periodic Table? *[Peaks represent the last element in a row of the table, the valleys represent the first element in a row.]*

c. Which elements have consistently high values and which ones are consistently low? *[Alkali metals are low, the noble gases are high.]*

d. How can these values (Question 3) be interpreted in terms of how strongly an atom holds electrons? *[It is easier to remove an electron from an alkali metal because it is not strongly held.]*

e. How would you predict the first ionization energy of cesium compares to that of sodium? (This question can be carried on to other families of elements, as well.) *[It should be easier to remove a valence electron, and therefore, have a smaller ionization energy.]*

f. How do the two irregularities in the data for the second and third periods relate to the electron arrangements of the element pairs involved? *[The dips correspond to an* s *to* p *electron and a half filled* p *to doubled up* p *configuration.]*

g. How would you expect a graph of second ionization energies to compare to your graph of first ionization energies? *[See section on second ionization energies.]*

Second Ionization Energies

a. Why was there no second ionization energy for hydrogen? *[Hydrogen only has one electron.]*

b. What similarities do you see between the first and second ionization energy graphs? *[A pattern of rising and falling peaks and valleys.]*

c. What differences do you see between the two graphs? *[The peaks represent the alkali metals.]*

d. How much higher are the second ionization energies than the first ionization energies? (Encourage students to make ratios of second to first ionization energies for the various elements.) *[They are 2-3 times as high except for the alkali metals that are ten times as high.]*

e. Which elements have highest second ionization energies? *[The alkali metals.]*

f. Are the irregularities of data within a period caused by electron arrangements similar to the ones that caused the irregularities in the first ionization energies? *[Yes; that is,* $s \rightarrow p$ *and* $p^3 \rightarrow p^4$*.]*

g. What explanation can be given for the huge jump in second ionization energies for the alkali metal elements? *[The second electron comes from a lower energy level.]*

h. What patterns would you expect to see in graphs of higher ionization energies – third, fourth, and fifth? *[Similar patterns, but with peaks moved over one element.]*

All ionization energy data are adapted from: Emsley, J. (1991). *The elements, 2nd ed.* Oxford:Clarendon Press. Values given in parentheses do not have a high degree of reliability.

$I_1 \quad M \rightarrow M^{+} + e^{-}$

$I_2 \quad M^{+} \rightarrow M^{2+} + e^{-}$

$I_3 \quad M^{2+} \rightarrow M^{3+} + e^{-}$

$I_4 \quad M^{3+} \rightarrow M^{4+} + e^{-}$

$I_5 \quad M^{4+} \rightarrow M^{5+} + e^{-}$

I_1 = First ionization energy, I_2 = second ionization energy, *etc.*

R_I = Relative ionization energies based upon the element in the list with lowest ionization energy (Rb). This choice will yield an ordinate ~100 units in length.

Ionization Energies and Relative Ionization Energies (based on Rb) for Selected Elements

Z	Symbol	I_1	RI_1	I_2	RI_2	I_3	RI_3	I_4	RI_4	I_5	RI_5
1	H	1312	3.26	-------		-------		-------		-------	
2	He	2372	5.89	5250	13.03	-------		-------		-------	
3	Li	513	1.27	7298	18.11	11815	29.32	-------		-------	
4	Be	899	2.23	1757	4.36	14848	36.84	21006	52.12	-------	
5	B	801	1.99	2427	6.02	3660	9.08	25025	62.10	32822	81.44
6	C	1086	2.69	2352	5.84	4620	11.46	6222	15.44	37827	93.86
7	N	1402	3.48	2856	7.09	4578	11.36	7475	18.55	9440	23.42
8	O	1314	3.26	3388	8.41	5300	13.15	7469	18.53	10989	27.27
9	F	1681	4.17	3374	8.37	6050	15.01	8408	20.86	11023	27.35
10	Ne	2081	5.16	3952	9.81	6122	15.19	9370	23.25	12177	30.22
11	Na	496	1.23	4562	11.32	6912	17.15	9543	23.68	13353	33.13
12	Mg	738	1.83	1451	3.60	7733	19.19	10540	26.15	13630	33.82
13	Al	577	1.43	1817	4.51	2745	3.14	11575	28.72	14839	36.82
14	Si	787	1.95	1577	3.91	3231	8.02	4356	10.81	16091	39.93
15	P	1012	2.51	1903	4.72	2912	7.23	4956	12.30	6273	15.57
16	S	1000	2.48	2251	5.59	3361	8.34	4564	11.33	7013	17.40
17	Cl	1251	3.10	2297	5.70	3826	9.49	5158	12.80	6540	16.23
18	Ar	1520	3.77	2665	6.61	3928	9.75	5770	14.32	7238	17.96
19	K	419	1.04	3051	7.57	4411	10.95	5877	14.58	7975	19.79
20	Ca	590	1.46	1145	2.84	4910	12.18	6474	16.06	8144	20.21
31	Ga	579	1.44	1979	4.91	2963	7.35	6200	15.38	(8700)	21.59
32	Ge	762	1.89	1537	3.81	3302	8.19	4410	10.94	9020	22.38
33	As	947	2.35	1798	4.46	2735	6.79	4837	12.00	6042	14.99
34	Se	941	2.33	2044	5.07	2974	7.38	4144	10.28	6590	16.35
35	Br	1140	2.83	2104	5.22	3500	8.68	4560	11.32	5760	14.29
36	Kr	1351	3.35	2350	5.83	3565	8.85	5070	12.58	6240	15.48
37	Rb	403	1.00	2632	6.53	3900	9.68	5080	12.61	6850	17.00
38	Sr	550	1.36	1064	2.64	4210	10.45	5500	13.65	6910	17.15
49	In	558	1.38	1821	4.52	2704	6.71	5200	12.90	(7400)	18.36
50	Sn	709	1.76	1412	3.50	2943	7.30	3930	9.75	6974	17.31
51	Sb	834	2.07	1794	4.45	2443	6.06	4260	10.57	5400	13.40
52	Te	869	2.16	1795	4.45	2698	6.69	3610	8.96	5668	14.06
53	I	1008	2.50	1846	4.58	3200	7.94	(4100)	10.17	(5000)	12.41
54	Xe	1170	2.90	2046	5.08	3097	7.68	(4300)	10.67	(5500)	13.65

4. Have students plot atomic radius *vs.* atomic number (from data table provided) and discuss patterns and trends. Data for this activity and suggestions for discussion follow.

   The number of data points in this graphing activity is limited. This is purposely done to force students to make generalizations from smaller amounts of information. If they have been through the analysis of ionization energy graphs, they have seen the way patterns repeat and have begun accounting for these patterns in terms of their understanding of atomic theory. This activity, as well as the one on ionic sizes, should enhance the learning from the ionization energy activity.

Graphs are plotted with atomic number on the horizontal axis, but the vertical axis will be atomic radius measured in picometers (1pm = 1 x 10^{-12} m). It is helpful to spread the atomic number coordinate so that the graph is larger and easier to interpret.

Suggestions for Discussion Questions

a. What trends and patterns do you see in the graph of these data?

b. Offer an explanation for the relationship between atomic number and atomic radius going from left to right in a period.

c. How would the atomic radii of lithium and beryllium compare to each other? Explain your answer.

d. How would the atomic radii of lithium and beryllium compare to the radii of the remaining elements of the second period? Explain your answer.

e. How would the sizes of the radii of the elements in the fourth period (K, Ca, Ga, *etc.*) compare to the sizes of the corresponding elements in the third period (Na, Mg, Al, *etc.*)? Offer an explanation for the differences in radii of the elements in these two periods.

Atomic and Ionic Radii for Some Representative Elements

Element	Atomic number (Z)	Atomic radius (pm)	Ionic charge	Ionic radius (pm)
Boron	5	88	3+	20
Carbon	6	77	4+	15
Nitrogen	7	70	3–	171
Oxygen	8	66	2–	140
Fluorine	9	64	1–	136
Sodium	11	157	1+	95
Magnesium	12	136	2+	65
Aluminum	13	125	3+	50
Silicon	14	117	4+	41
Phosphorus	15	110	3–	212
Sulfur	16	104	2–	184
Chlorine	17	99	1–	181

f. Have students analyze and discuss ionic size related to family, ionic charge, and atomic numbering, deriving information on ionic sizes from the data table provided. Data for this activity and suggestions for discussion follow.

5. This activity is performed similarly to the atomic radii activity, using the same elements' ions. The same questions from the other activity can be adapted to fit this activity.

6. Have students generate electron configurations for selected families and/or periods of the Periodic Table and discuss similarities and/or trends in the electronic structures. Relate and justify these trends with physical properties. There must be hundreds of examples of alternative periodic tables, some humorous and others serious. There are also periodic tables used in the textbooks of countries that don't use an English alphabet. These are interesting and make good displays. A few examples of alternative periodic tables are included here.

7. The first ionization energy data (see *Other Activities 4*) can be used to construct three dimensional graphs that illustrate periodic trends. Glue a periodic table to a board and drill small holes through each box representing the elements. Use medical swab sticks cut to scale into lengths representing the relative ionization energies. A 96-well plate can also be used. Straws that fit into the holes can be cut to scale. The periodic trends of other chemical and physical properties can be explored in this manner.

8. The "Great Periodic Table Race" is a board game using dice, game pieces, and "discovery cards." Game was developed by Richard Rodin and distributed by Science Kit.

9. "Understanding the Periodic Table" is an analogy puzzle published by Hubbard Scientific Co., 1952-T Raymond Drive, Northbrook, IL 60620. (708) 272-7810 or FAX (708) 272-9894.

10. Have students generate a concept list during a class brainstorming session at the end of the discussion of periodicity. Put concepts on cards and generate a map (in groups of two). In groups of four, have students develop a consensus map. Develop a consensus (total class) map during a teacher led discussion.

11. Many types of alternative periodic charts have been proposed. Obtain copies of these charts and discuss the pros and cons of the format of the elements.

    Five alternatives have appeared together on a poster [*Chemistry Currents,* (1). Fall 1986—CentCom, Ltd., 60 East 42nd Street, New York, NY 10165; (212) 976-9660.] and separately in *Journal of Chemical Education.*

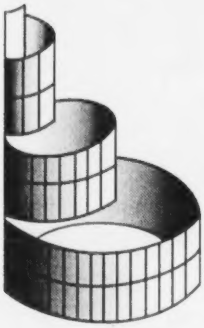
    a. Octagonal prismatic periodic table [Kow, T. W. (1972). *49,* 59].
    b. Pyramidal format [Fernelius, W. C., and Powell, W. H. (1982). *59,* 504].
    c. Three-dimensional form consisting of two concentric helices [Quam, G. N., and Quam, M. B. (1934). *11,* 288].
    d. Fan-shaped table [Quam, G. N., and Quam, M. B. (1934). *11,* 288].
    e. Circular format [Strong, F. C. III. (1985). *62,* 456].

    Three alternatives (two reproduced in the *Appendix*) have appeared on calendars published by Instruments for Research and Industry (I^2R), Inc., 108 Franklin Avenue, Box 159, Cheltenham, PA 19012.

    f. Spiral Periodic Chart (September 1982)
    g. Periodic Pyramid (September 1984)
    h. 3-D Periodic Chart (September 1983)

12. Construct your own Periodic Spiral.

Several years before Mendeleev's Table, A.-É. Beguyer de Chancourtois published the first spiral periodic table in 1863. Spiral periodic tables are especially useful for showing the continuity of the elements, as well as the relationships between the *main group elements*, the *transition elements*, and the *lanthanides* and *actinides*.



This spiral three-dimensional table is based on a 1914 design by Ingo Hackh. The flat periodic chart in the *Appendix* should be used to construct the spiral version. Copy on heavy stock paper. Use a coloring pencil and color the main group elements blue; the transition elements, orange; and the lanthanides and actinides, brown. To assemble your spiral periodic table, cut along the heavy black lines. With the printing facing out, bend into three circles of different diameters. Position the elements Li, Na, K, Rb, Cs, and Fr so that they overlap the corresponding symbols on the tab. Glue in place.

Figure 4. Periodic spiral.

TIPS FOR THE TEACHER

Language of Chemistry

alkali metals elements in group headed by Li.

alkaline earth metals elements in group headed by Be.

atomic radius half the distance of a single bond between two atoms of an element.

electronegativity a measure of attraction of an atom for a pair of shared electrons.

first ionization energy energy required to remove the most loosely held electron from the outer energy level of an atom in the gas phase.

group or **family** a vertical column in the Periodic Table.

halogens elements in group headed by F.

noble gases elements in group headed by He.

period horizontal row in the Periodic Table.

periodic law chemical and physical properties of elements are periodic functions of their atomic number.

transition element any *d*-block element of the Periodic Table.

Common Student Misconceptions

1. **"The radius of an atom increases in a period with the addition of electrons."**

From the data given on atomic radii, we see that with the addition of electrons, the atomic radii decrease within a period. This is due to the increase in the positive charge of the nucleus and its attraction for the electrons in the atom. Although the number of electrons in the atom increases as one goes across a period and although the electrons repel each other potentially causing the atoms to grow in size, this effect is more than opposed by the increase in the number of protons attracting these electrons. Since these electrons are added to the same shell, the number of protons attracting them causes the size of the atoms to decrease.

2. **"All properties show periodic characteristics."**

   There are deviations in all periods.

3. **"Elements are arranged according to atomic weight."**

   While this is true in general, the periodic properties are characteristic of atomic numbers.

HISTORY: ON THE HUMAN SIDE

Origins of the Periodic Table are rooted in the idea of elementary substances first proposed by the ancient Greeks. In 450 B.C., Empedocles proposed the existence of four elementary substances: earth, air, fire, and water. These were later adopted by Plato and Aristotle (who added a fifth, *quinta essentia*, crystalline spheres of the heavens). The history of elements paralleled that of atoms until 1700 when atoms were becoming well established while the idea of four elements was questioned. A clearly defined working concept of elements was proposed by John Dalton in 1808. In 1817, Johann Wolfgang Döbereiner pointed out that elements could be grouped in triads of related substances (Li-Na-K; Cl-Br-I; Ca-Sr-Ba; S-Se-Te; As-Sb-Bi). He found that when these triads of elements were arranged in order of increasing mass, a series emerges in which the intermediate atomic mass was close to the arithmetic average of the other two. Properties such as density, melting point, and boiling point behaved similarly.

The time was ripe for a classification according to similarities. Taxonomy was beginning to be a common practice in biology, and it was felt that chemistry also needed a taxonomy. What was lacking was a set of reliable atomic masses on which to base the taxonomy. This fact was established in 1860 at an international conference in Karlsruhe, Germany. Both Dmitri Mendeleev and Lothar Meyer attended this conference.

Alexandre-Émile Beguyer de Chancourtois (1820-1886), a French geologist, arranged elements in a three dimensional spiral in order of increasing atomic mass. Elements with similar properties fell into vertical columns. John Newlands (1827-1898) proposed a law of octaves when he noticed that every known eighth element, in order of increasing atomic mass, repeated properties. The stage was set for Mendeleev and Meyer.

Dmitri Mendeleev (1834-1907) was a Russian born in Siberia, the youngest of 13, 14, 17, or 18 children depending upon the historical source. When his father died, his mother took him on a 1000 mile journey by horse cart to Moscow and then to St. Petersburg, where he was educated. His hot temper led him to quarrel with an important minister of education, and he was unable to obtain a university position. As a consequence, he was briefly a high school teacher and then after continuing his education, did research in Paris and Heidelberg. Returning to Russia, he was appointed Professor of Chemistry at the Technological Institute of St. Petersburg.

In 1869, Mendeleev was writing his textbook, *Principles of Chemistry*. As part of the project he was searching for a way of organizing the vast amount of information concerning the elements as an aid to his students' understanding. On February 17, 1869, he was arranging a stack of cards on which he had written the name, symbol, and atomic mass of each element. He laid out the cards in rows in order of increasing mass and noticed what others before him had—the elements fall into families. He published his table listing 63 known elements, leaving spaces for 31 more. Mendeleev not only predicted the existence of unknown elements; he predicted their properties. His most successful predictions were for the elements scandium, gallium, and germanium, which were later discovered. In addition, he suggested corrections for atomic masses on purely theoretical grounds; these suggestions also later proved to be accurate. These predictions showed the power of his periodic table.

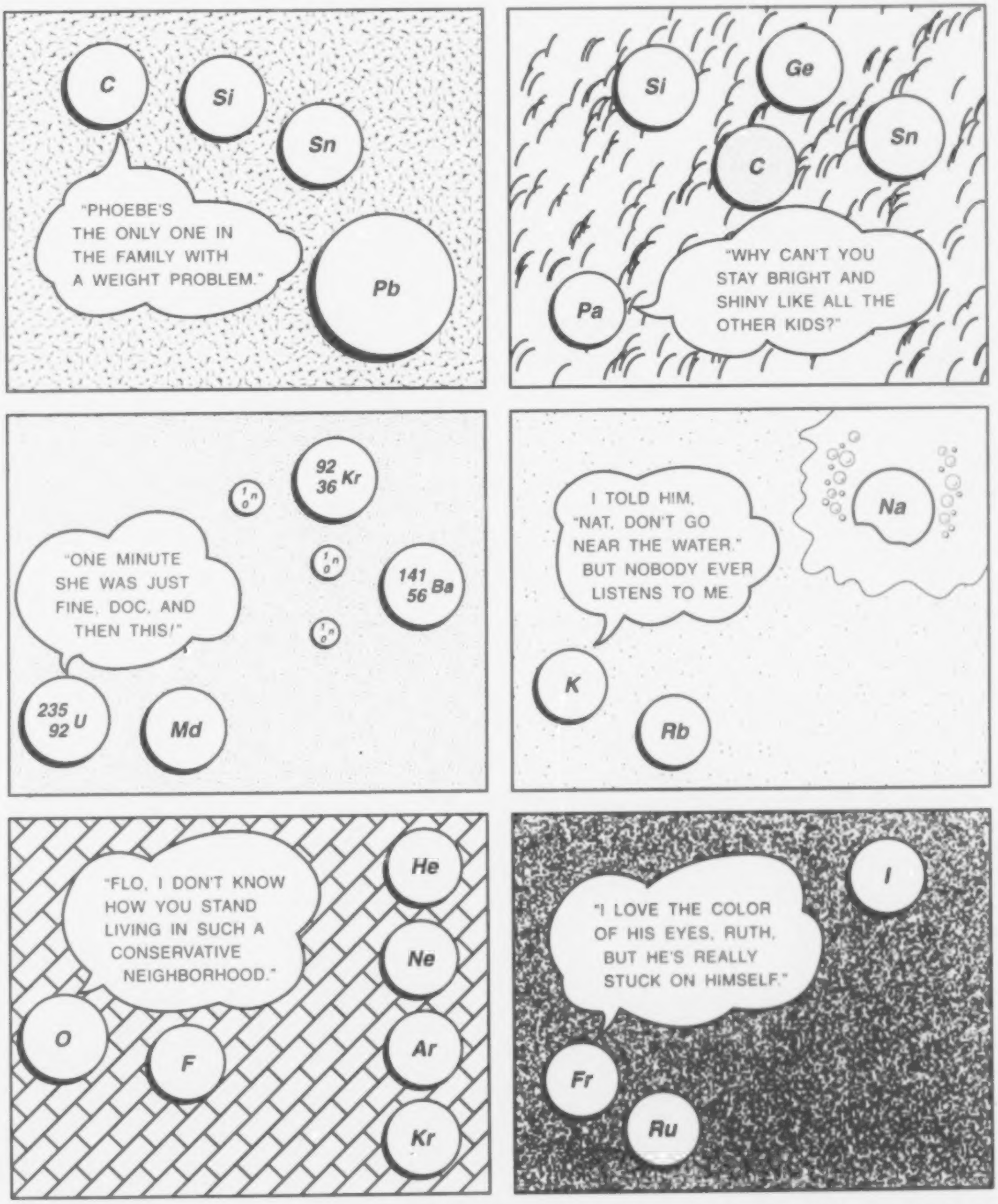
Mendeleev had a hot temper and strong democratic leanings, which got him into trouble in Czarist Russia. For the most part, he was tolerated by his country's officials because of his fame and importance as an internationally known chemist. However, in 1890, he was forced to resign his university position in a policy dispute with the administration over oppression of students. At this time, he took over as Director of the Bureau of Weights and Measures.

Lothar Meyer (1830-1895), a German chemist, proposed an arrangement independent of and similar to that of Mendeleev. His arrangement emphasized the periodicity of physical properties. He made graphs of atomic volume versus atomic weights that showed a line which rose and fell into cycles or periods with minima and maxima defining chemical families. Meyer, like Mendeleev, made his proposal as a result of writing a textbook. He did not publish his findings until 1870, one year after Mendeleev had. He and Mendeleev engaged in an unfortunate quarrel over priority. He later admitted that he lacked the courage to predict the existence of undiscovered elements as Mendeleev had.

Daniel Morris (see *References*) suggests a connection between music and the development of ideas of elements.

HUMOR: ON THE FUN SIDE

1. What elements think of each other, a conversation between elements. Ihde, J. (1986, February). Puzzle page. *Chem Matters*, p. 20. Washington: American Chemical Society.


C
Si
Sn
"PHOEBE'S THE ONLY ONE IN THE FAMILY WITH A WEIGHT PROBLEM."
Pb
Si
Ge
C
Sn
Pa
"WHY CAN'T YOU STAY BRIGHT AND SHINY LIKE ALL THE OTHER KIDS?"
$^{1}_{0}n$
$^{92}_{36}Kr$
$^{1}_{0}n$
$^{141}_{56}Ba$
$^{1}_{0}n$
"ONE MINUTE SHE WAS JUST FINE, DOC, AND THEN THIS!"
$^{235}_{92}U$
Md
I TOLD HIM, "NAT, DON'T GO NEAR THE WATER." BUT NOBODY EVER LISTENS TO ME.
Na
K
Rb
"FLO, I DON'T KNOW HOW YOU STAND LIVING IN SUCH A CONSERVATIVE NEIGHBORHOOD."
He
Ne
Ar
Kr
O
F
I
"I LOVE THE COLOR OF HIS EYES, RUTH, BUT HE'S REALLY STUCK ON HIMSELF."
Fr
Ru

John Ihde, chemistry teacher at Wausau West High School, Wausau, WI captured some rare conversation between the elements.

2. **Fantastic New Compounds**. Clues given on elements to make compounds. Russell, F., *et al.* (1986, October). Puzzle page. *Chem Matters*, p. 1, 16. Washington: American Chemical Society.

ChemMatters magazine has obtained a bizarre research report describing a previously unknown class of compounds made by a new principle of chemical synthesis. One newly discovered compound, which contains the elements *indium* (In) and *potassium* (K), can be used in ballpoint pens. Remarkably, its formula is *InK*! The formulas of the other compounds were not revealed, but can be derived from their chemical composition. Can you write the formulas?

a. A pen to hold the ink can be made of a compound containing one part bismuth and one part carbon.

b. A compound with equal parts tin, oxygen, and boron does not mix with other substances.

c. One part phosphorus and nine parts potassium make up this compound that was recently isolated from dog urine.

d. A very absorbent material of equal parts sulfur, phosphorus, oxygen, nitrogen, and germanium has multiple uses.

e. A mineral of the three elements aluminum, sulfur, and phosphorus is found in a major mountain range.

f. The compound containing lithium, phosphorus, and sulfur forms a structure that enables you to smooch.

g. This chemical, sometimes found in rabbits' feet, contains one part lutetium, one part carbon, and one part potassium.

h. A compound that's helpful when you're studying for a test contains one part chromium and one part americium.

i. Bromine, iodine, carbon, and potassium are found in this material that is used to build houses.

j. This compound, which was banned by the church, contains one part nitrogen and one part silicon.

k. Discovered in automobiles, this compound contains one part sulfur and one part gallium.

l. A substance one part boron, one part oxygen, and one part astatine floats on water.

m. This compound—one part aluminum, one part molybdenum, and one part neodymium—makes a very nutty substance.

n. A compound of one part potassium, one part nickel, and one part iron penetrates many other substances.

o. Tungsten, rhenium, nitrogen, carbon, and hydrogen in equal parts can form a useful tool.

p. Equal parts of carbon, oxygen, calcium, cobalt, and lanthanum will make a tasty beverage.

q. A substance that's one part cobalt, one part radon, and one part yttrium is similar to many of these examples.

Answers:

1. BIC 2. SNOB 3. K_9P 4. SPONGE 5. ALPS 6. LIPS 7. LUCK 8. CRAM 9. BRICK 10. SIN 11. GAS 12. BOAT 13. ALMOND 14. KNIFE 15. WRENCH 16. COCA COLA 17. CORNY

3. **Periodic Puns**

Some very uncommon elements (see *Appendix*) from Huheey, J. (1978). *Diversity and periodicity: An inorganic chemistry module.* New York: Harper & Row.

a. When people leave, they ___. *[Argon]*

b. What you do with dead people. *[Barium]*

c. Old McDonald put the sow in the field and the ___ the porch. Eeii, Eeii, Ooo. *[Boron]*

d. One who follows a golfer around the course. *[Cadmium]*

e. An early morning "snap, crackle, and pop!" *[Cerium]*

f. A salt water chemical. *[Cesium]*

g. People are always doing this to their hair on a windy day. *[Chromium]*

h. What country is known for Paris, perfume and profound leaders? *[Francium]*

i. How does one purchase gasoline (by the ___)? *[Gallium]*

j. "On your mark...Get set...___!" *[Gold]*

k. Romeo fell head over ___ with Juliet. *[Helium]*

l. "Oh! Give me a ___ where the buffalo roam..." *[Holmium]*

m. Water and gin combo. *[Hydrogen]*

n. What you do with wrinkled clothes. *[Iron]*

o. A dead person lies in a ___ a hill in the cemetery. *[Krypton]*

p. To scrub the floor, you must get down and put your ___ it. *[Neon]*

q. Who was the main character in the television show "Have Gun Will Travel"? *[Paladium]*

r. ___ your boots before you to go work and take them off when you get home. *[Polonium]*

s. Those who hold wild parties find that the police will eventually ___. *[Radium]*

t. A cool dude might exclaim this in response to a perfect test score. *[Radon]*

u. To answer these awful puns is to ___! *[Sulfur]*

v. A person in a bank who deposits your money for you. *[Tellurium]*

w. An apologetic lisping child says this. *[Thorium]*

x. A truck driver's farewell (___ –4 Good Buddy!) *[Tin]*

y. If you yawn near a beehive, you might get your ___. *[Tungsten]*

z. A boat with a hole in the bottom will ___. *[Zinc]*

4. Tom Lehrer's *An Evening Wasted with Tom Lehrer* contains a cut called "Element Song" in which the element names are sung to Sir Arthur Sullivan's "Mighty Major General." Available on CD (Warner #6199). Contact local tape/CD store.

5. **Cartoon**. Knights of the Periodic Table. Sources: Siromoski, R. *Discover*, May 1990, p. 93 (see *Appendix* for cartoon).

6. A Dear John letter from Francium. Source: Ratzlaff, B. (December 1985). *ChemMatters*, p. 13.

Dear John,

I'm writing to tell you that I can't marry you—I'm breaking our engagement. I guess you want to know why, so let me explain. I'll start at the beginning...

It all started when I was born the daughter of Actinium-227 by alpha emission. I not only grew up unstable, but my psychiatrist recently diagnosed me as a paranoid schizophrenic and said she has trouble telling my 20 isotopes apart. The sad truth is that ^{223}*Fr (my longest lived isotope) has a half-life of only 21 min, after which I decay into that awful* ^{223}*Ra! And my* ^{221}*Fr has a half-life of only 4.8 min.! So you see, there's no use continuing our relationship when I won't be around long.*

I was a quiet child. When my existence was finally discovered in 1939 by Marguerite Perey at the Curie Institute in Paris, I was thrilled! They named me Francium after the country in which I was discovered. They introduced me to my sisters: Lithium, Sodium, Potassium, Rubidium, and Cesium. Soon I learned that, as members of the Alkali metal family, we had a lot in common. We all had a valence of one, tarnished in air, had low melting points, and reacted with water. Not only that, we generally had soft crystals and were commonly found as halides and as aluminosilicates and combined vigorously with other elements.

Father said that the strong similarity between my sisters and me was the arrangement of the electrons in our atoms. But then, Father always had an explanation for everything. I asked him once why he liked Lithium best. He said he loved us all equally and that I was only being my unstable self. But I keep bugging him, and finally he admitted that he liked Lithium best because she was used in the treatment of steel parts and was making something of her life. He said I had no meaningful purpose that he could see. Then he looked me up and down and grunted that I ought to do something about my atomic weight. I ran to my room in tears and looked in the mirror. My atomic weight was around 223, more than anyone else in my Alkali metal family. John, I just can't bear it! You saw how futile my attempts to diet were.

Sitting in the earth's crust the way I do in tiny amounts (never more than one ounce), I have a lot of time to think. It isn't so bad having a melting point of 27 °C, and my changes of phase add some excitement to my life. But something has been bothering me. Even though I was discovered more than 30 years ago, my sisters still leave me out of everything. The Alkali family has always been famous, but nobody knows the real *me! Why, I remember that in* Inorganic Chemistry, *R.T. Sanderson said, 'Relatively little is known of this element except that a close resemblance to Cesium has been recognized.'*

Cesium! He compared me with Cesium—my sister who hangs out in a mineral called pollucite. No one would ever catch a weighable amount of me *in that trash! I'm sorry—there I go being unstable again.*

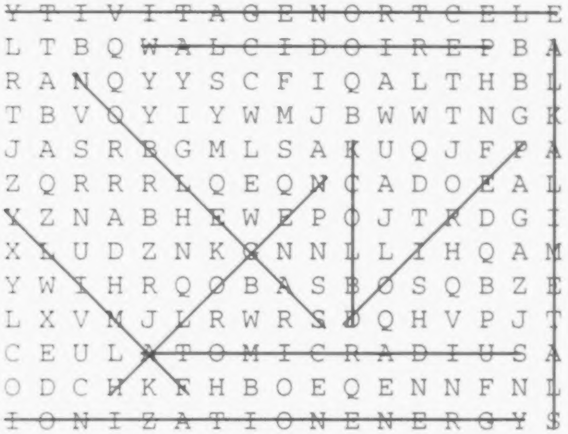
Hey, did you hear that people are actually cloning me? It's true! What they do is bombard thorium with protons, and they've got instant artificially made Francium. Neat, isn't it? Who knows? Maybe some day, good old atomic number 87 will find her niche in society, and I will be accepted for what I am! But until then, I know we could never be right for each other.

John, I am really sorry. Please try to understand. I will always love you and will never forget our half-lives together.

Love always,

Francium Alkali

7. Word Search (see *Appendix* for master copy)



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Y T I V I T A G E N O R T C E L E
L T B Q W A L C I D O I R E P B A
R A N Q Y Y S C F I Q A L T H B L
T B V O Y I Y W M J B W W T N G K
J A S R B G M L S A K U Q J F P A
Z Q R R R L Q E Q N C A D O E A L
Y Z N A B H E W E P O J T R D G I
X L U D Z N K G N N L L I H Q A M
Y W I H R Q O B A S B O S Q B Z E
L X V M J L R W R S D Q H V P J T
C E U L A T O M I C R A D I U S A
O D C H K F H B O E Q E N N F N L
I O N I Z A T I O N E N E R G Y S
```

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Half the distance of a single bond between two atoms of an element (2 words).
2. Measure of the attraction of an atom for a pair of shared electrons.
3. Energy required to remove the most loosely held electron of an atom in the gas phase (2 words).
4. Helium is one (2 words).
5. Astatine is one.
6. Elements in the first group (2 words).
7. Another name for group.
8. A horizontal row on the periodic chart.
9. Transition elements are in this section of the periodic chart.
10. ______ states that chemical and physical properties of the elements are periodic functions of their atomic number (2 words).

Answers: 1. ATOMIC RADIUS 2. ELECTRONEGATIVITY 3. IONIZATION ENERGY 4. NOBLE GAS 5. HALOGEN 6. ALKALI METALS 7. FAMILY 8. PERIOD 9. D-BLOCK 10. PERIODIC LAW

MEDIA

1. The *World of Chemistry* videotape "Number 7: The Periodic Table." World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224. A secondary school version of this series is available from WINGS for Learning/SUNBURST, 101 Castleton Street, Pleasantville, NY 10570; (914) 747-3310; (800) 321-7511; (914) 747-4109 (FAX).

2. Other films and videos.

   a. *The Periodic Table*. Presents the history, terminology and structure of the Periodic Table and explains the relationship between electron configuration and preiodicity. 24 min. video. The Media Guild, 11526 Sorrento Valley Road, Suite J, San Diego, CA 92121; (619) 755-9191; (800) 886-9191.

   b. *Chemical Families*, A CHEM Study Film. Ward's Natural Science Establishment, Inc., 5100 West Henrietta Road, P.O. Box 92912, Rochester, NY 14692-9012; (716) 359-2502; (800) 962-2660; (716) 334-6174 (FAX).

   c. *Ascent of Man Part 10: World Within World*, KCET Video, 4401 Sunset Blvd., Los Angeles, CA 90027; (800) 328-7271.

   d. *Updating the Periodic Table*. Videotape address by Glenn Seaborg. Item Number DC 5800. Lawrence Hall of Science Discovery Corner, University of California at Berkeley, Berkeley, CA 94720; (510) 642-1929.

   e. "Periodic Classification of the Elements" and "Periodic Properties of the Elements," two chapters from the Chem 101 Series, Learning Center, Department of Chemistry, University of Illinois at Urbana-Champaign, Urbana, IL 61801; (217) 333-2998.

3. Software published by*JCE:* Software, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   a. *KC? Discoverer with Knowledgeable Counselor,* by Daniel Cabrol, John W. Moore and Robert C. Rittenhouse. Special Issue 2, for IBM PS/2, PC compatible computers.

   b. *KC? Discoverer: Exploring the Properties of the Chemical Elements,* by Aw Feng and John W. Moore. Vol. I B, No. 1, for IBM PS/2, PC compatible computers.

   c. *KC? Discoverer?,* by Michael Liebl, Vol. IV A, No. 2, for all Apple II computers.

   d. *The Periodic Table Stack,* by Michael Farris. Vol. I C, No. 1, for the Apple Macintosh.

   e. *The Periodic Table* (Toolbook), by Paul F. Schatz, John C. Kotz and John W. Moore, in press. For Windows running on IBM PS/2 and PC-compatible computers.

   f. *Aufbau*, by Dale Jensen, one of the programs in *The One-Computer Classroom: Periodic TableWorks*, Vol. I A, No. 1 for the Apple II computer.

4. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).

For the Apple II computer:

AP 207, *The Periodic Table as a Database*. The Periodic Table is used as a graphical database to display chemical and physical properties. Various display modes, including 3-D, are accessible.

AP 209, AP 310

5. Other Software.

   a. *Astro Stack* HC17. Each element has its own valuable information. HyperCard program for Macintosh. The Public Domain Exchange, 2074C Walsh Avenue, Dept. 644, Santa Clara, CA 95050; (408) 955-0292.

   b. *MacMendeleev*. A computerized Periodic Table database for the Macintosh. Trinity Software, P.O. Box 960, Campton, NH 03223; (603) 726-4641.

6. Videodiscs published by *JCE:* Software, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, Wl 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   a. "Electron Clouds" and "The Aufbau Principle," two chapters on The World of Chemistry: Selected Demonstrations and Animations: Disc II (double sided, 60 min.), Special Issue 4.

   b. "Similarities and Trends in Groups I – Noble Gases" and "Similarities and Trends in Groups II – Alkali Metals," two chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc I (double sided, 60 min.), Special Issue 3.

   c. *The Periodic Table Videodisc* (single side, 30 min.), Special Issue 1.

Links/Connections

WITHIN CHEMISTRY

1. Periodicity is a central organizing feature of chemistry. Consequently, it is linked to many major concepts in chemistry—electronic structure, atomic structure, and bonding. It is related to the atomic properties of elements (*e.g.*, electronegativity) and to their empirical properties (*e.g.*, reactivity).
2. The spectra of a family of elements such as the alkali metals can be correlated with electron transitions from the excited states to lower energy states. The collective transitions give rise to characteristic colored flames, lithium (red), sodium (yellow), potassium (lilac), rubidium (purple), and cesium (blue). The brilliant display of color in fireworks, flares, and aerial rockets takes place when chemical changes supply sufficient thermal energy to the salts of alkali metals so that the metal ions are promoted to excited states, and radiate the energy as characteristic colors during the relaxation process to lower energy states.

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

The Periodic Table classifies and organizes knowledge in chemistry. As such, it is parallel to classification systems in other disciplines—for example, to taxonomic classification in biology and Library of Congress or Dewey Decimal systems in libraries (see *Metaphors and Analogies*).

TO THE CONTEMPORARY WORLD

Personal

Students may relate some elements of the Periodic Table to everyday life. Examples are:

1. **Fluoridation**. Fluorine is beneficial in preventing cavities.
2. **Calcium**. Calcium is a mineral needed for strong bones and teeth.
3. **Lead**. The issue of lead poisoning.
4. **Aluminum**. The issue of recycling resources.
5. **Chlorine**. The use of chlorine in water to prevent disease.
6. **Helium**. Use in balloons. Non-toxic but may asphyxiate. Sources: Natural gas may contain as much as 7% helium. Can you give some reasons why? What might be the origin of He in the earth's crust? Relate to *Nuclear Chemistry* module.
7. **Germanium and Silicon**. Role in semiconductors.
8. **Titanium**. Use in "white out" materials such as Liquid Paper™. Use in paints.

Community

1. **Field trips**. Mining or manufacturing involving elemental substances in your area.
2. **Knowledgeable individuals**. An M.D. with an interest in environmental or public health might be an excellent resource to discuss the presence of elemental substances such as mercury or lead in our environment.

3. The prevalence of lead poisoning among inner-city children. Causes, prevention; chemical antidotes for lead poisoning.
4. **Nitrogen**. Role as a fertilizer; liquid NH_3; nitrogen fixation.
5. **Role of metals in society**. Metallurgy. Extraction of metals from their ores. Alloys. The role of steel in our lives. Large structures demanding great strength would not be possible without the use of steel (*e.g.*, skyscrapers and superhighways).
6. **Zinc**. Galvanization (placing a protective coating of zinc on iron). Use of zinc in dry cells (ordinary flashlight batteries). Use of zinc oxide as a topical ointment.
7. **Tin**. Uses in tin plate for canned foods. Also used in many important alloys, including solder and pewter.
8. Role of certain elements in materials science (see *Materials Science* module).
9. **Transuranium elements**. Can we predict the properties of elements that have not yet been synthesized? What role do some of the transuranium elements play in our lives?

Extensions

1. Have students do written and/or oral reports on the discovery, physical and chemical properties and common uses of selected elements.
2. Have students obtain data on ionization energies, atomic sizes and ionic sizes for the transition elements in Periods 4 and 5. Discuss patterns (or lack of patterns) relating to these data. (Data are given in *Appendix* of *Transition Elements* module.)
3. Have students explore density data for elements to see if density could be considered a periodic property.
4. Have students obtain electronegativity data for elements and graph them to ascertain periodic trends for this property.
5. Primo Levi's collection of short stories *The Periodic Table* (see *References*) provides good extension topics for discussion. Have different students read the various stories and give book reports on them.

References

Module drafted by Michael Abraham, Donna Coshow, and William Fix, the Oklahoma team.

Ball, D. W. (1985). Elemental etymology: What's in a name? *Journal of Chemical Education, 62*(9), 787-788.

How the elements were named.

Brooks, D. W. (Producer). (1989). Doing chemistry [Videodiscs, computer program, and supporting written materials]. Washington, DC: American Chemical Society.

Cotton, F. A., and Wilkinson, G. (1988). *Advanced inorganic chemistry* (5th Ed.). Wiley: New York.

An advanced college text filled with descriptive information concerning the elements and their compounds.

Emsley, J. (1991). *The elements* (2nd Ed.). Oxford: Clarendon Press.

A comprehensive compilation of data on the elements. Very handy reference book.

Fernelius, W. C., and Powell, W. H. (1982). Some confusion in the Periodic Table of the elements. *Journal of Chemical Education, 59,* 504-508.

A historical review of the group numbering system in the Periodic Table.

Holden, N. E. (1984). *A history of the Periodic Table: SIRS science series.* Boca Raton, FL: Social Issues Resources Series, Inc.

This will augment your knowledge of historical information concerning the Periodic Table.

Levi, P. (1984). *The Periodic Table.* New York, NY: Random-Pantheon-Schocken.

A work of fiction with essays, short stories, and biographical remembrances organized around elements in the Periodic Table. This book was written by a distinguished Italian chemist/writer.

Loening, K. L. (1984). Recommended format for the Periodic Table of the elements. *Journal of Chemical Education, 61,* 136.

A note on the ACS-recommended format for the Periodic Table.

Morris, D. (1969). Music of the new spheres. *Chemistry, 42*(11), 10-12.

The author suggests a connection between music and the development of ideas of elements.

Nelson, B. J. (1984, April). Elements and their organization. *Chem Matters.*

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Orna, M. V. (1982). On naming the elements with atomic number greater than 100. *Journal of Chemical Education, 59,* 123.

A review of the origin and meaning of the three-letter symbols of the elements above atomic number 100.

Partners for a noble element. (1987, November 21). *Science News.*

Roberts, R. M. (1989). *Serendipity: Accidental discoveries in science*. New York, NY: Wiley.

Chapter 7 of this book is on "Discoveries of Chemical Elements" and discusses oxygen, iodine, helium, and the noble gases.

Rochow, E. G. (1977). *Modern descriptive chemistry*. Philadelphia, PA: Saunders.

More details than can be obtained in most beginning college chemistry texts.

Sanderson, R. T. (1960). *Chemical periodicity*. New York, NY: Reinholt.

A classic text on periodicity.

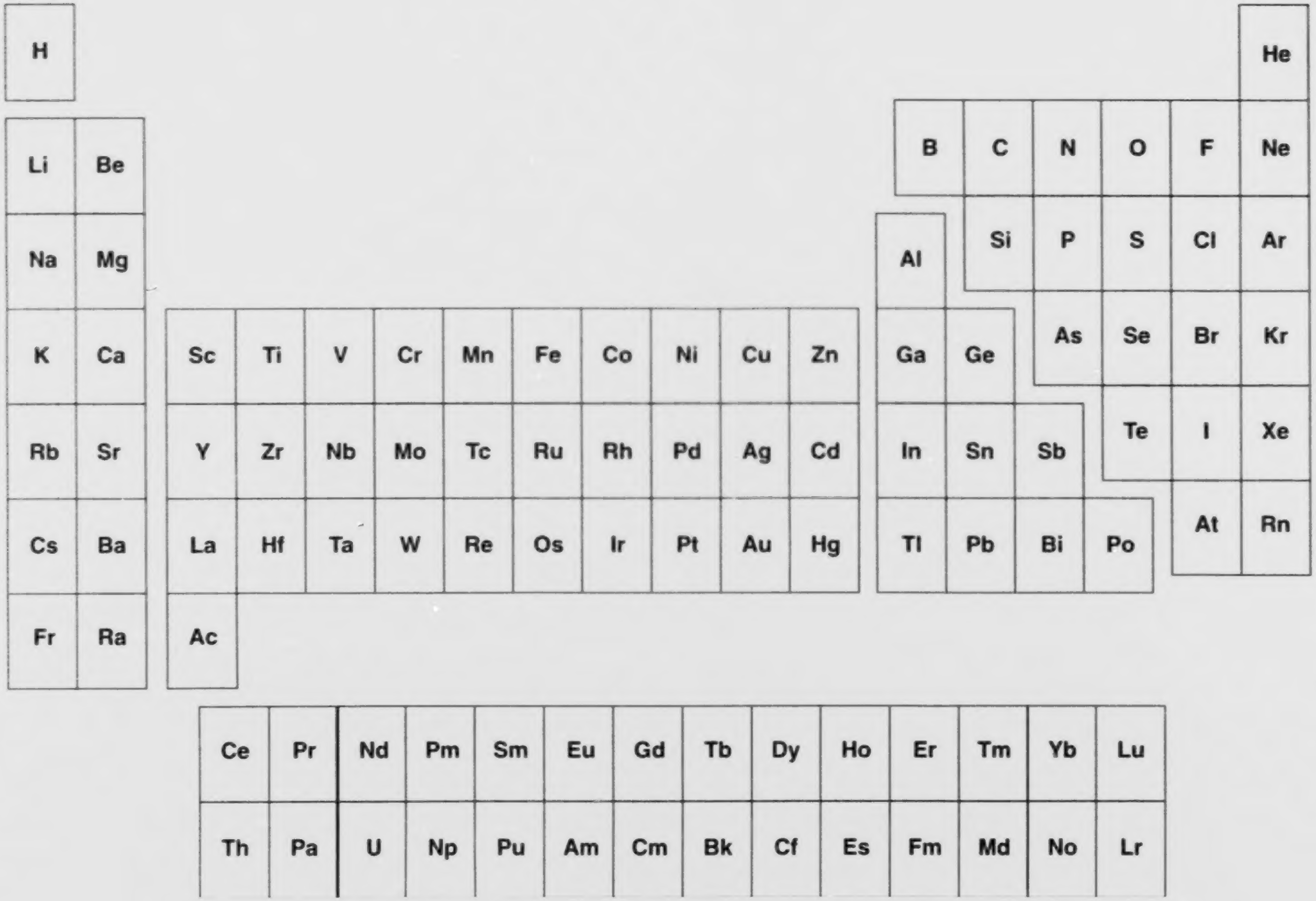
Weast, R. C. (Ed.). (1992). *Handbook of chemistry and physics* (any edition). Cleveland, OH: CRC Press.

The source of data about everything chemical including the elements and their compounds.

Appendix

- **Transparency Masters**
  1. Divided Periodic Table
  2. Densities of the Elements (see text by Sanderson in *References*)
  3. Melting Points of the Elements (see text by Sanderson in *References*)
  4. Word Search
- **Tables**
  1. Identification of Selected Substances
- **Other**
  1. I^2R Periodic Chart: The Periodic Pyramid
  2. I^2R Periodic chart: Non Scientist's Concept of the Periodic Chart of the Elements
  3. Spiral Periodic Chart Cut-Out
- **Humor**

DIVIDED PERIODIC TABLE



H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac															
			Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	
			Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr	

DENSITIES OF THE ELEMENTS

(g/cm^3) or (g/L) with *
gr = graphite
w = white
g = gas
l = liquid

1	*H (g) 0.09																*H (g) 0.09	*He (g) 0.18
2	Li 0.53	Be 1.85											B 2.4	C (gr) 2.26	*N (g) 1.25	*O (g) 1.43	*F (g) 1.69	*Ne (g) 0.90
3	Na 0.97	Mg 1.74											Al 2.70	Si 2.3	P (w) 1.82	S 2.07	*Cl (g) 3.2	*Ar (g) 1.78
4	K 0.86	Ca 1.55	Sc 2.99	Ti 4.5	V 5.96	Cr 7.1	Mn 7.2	Fe 7.86	Co 8.9	Ni 8.90	Cu 8.92	Zn 7.14	Ga 5.91	Ge 5.36	As 5.7	Se 4.8	Br (l) 3.11	*Kr (g) 3.74
5	Rb 1.53	Sr 2.6	Y 5.51	Zr 6.5	Nb 8.6	Mo 10.2	Tc 11.5	Ru 12.2	Rh 12.5	Pd 11.4	Ag 10.5	Cd 8.6	In 7.3	Sn 7.28	Sb 6.7	Te 6.2	I 4.93	*Xe (g) 5.89
6	Cs 1.90	Ba 3.5	La 6.15	Hf 13.3	Ta 16.6	W 19.3	Re 20.5	Os 22.48	Ir 22.4	Pt 21.45	Au 19.3	Hg (l) 13.55	Tl 11.85	Pb 11.34	Bi 9.8	Po 9.4	At	*Rn (g) 9.73

MELTING POINTS OF THE ELEMENTS

(°C)

1	● H -259.2																● H -259.2	● He -269.7
2	● Li 180.5	● Be 1283											● B 2027	● C	● N -210.1	● O -218.8	● F -219.6	● Ne -248.6
3	● Na 98	● Mg 650											● Al 660	● Si 1423	● P 44.2	● S 119	● Cl -101	● Ar -189.4
4	● K 63.2	● Ca 850	● Sc 1423	● Ti 1677	● V 1917	● Cr 1900	● Mn 1244	● Fe 1539	● Co 1495	● Ni 1455	● Cu 1083	● Zn 419.5	● Ga 29.78	● Ge 960	● As 817	● Se 217.4	● Br -7.2	● Kr -157.2
5	● Rb 38.8	● Sr 770	● Y 1500	● Zr 1852	● Nb 2487	● Mo 2610	● Tc 2127	● Ru 2427	● Rh 1966	● Pd 1550	● Ag 960.8	● Cd 320.9	● In 156.2	● Sn 231.9	● Sb 630.5	● Te 450	● I 113.6	● Xe -111.9
6	● Cs 28.6	● Ba 710	● La 920	● Hf 2222	● Ta 2997	● W 3380	● Re 3180	● Os 2727	● Ir 2454	● Pt 1769.3	● Au 1063	● Hg -38.87	● Tl 303.6	● Pb 327.4	● Bi 271.3	● Po 254	At	● Rn -71

197

Word Search

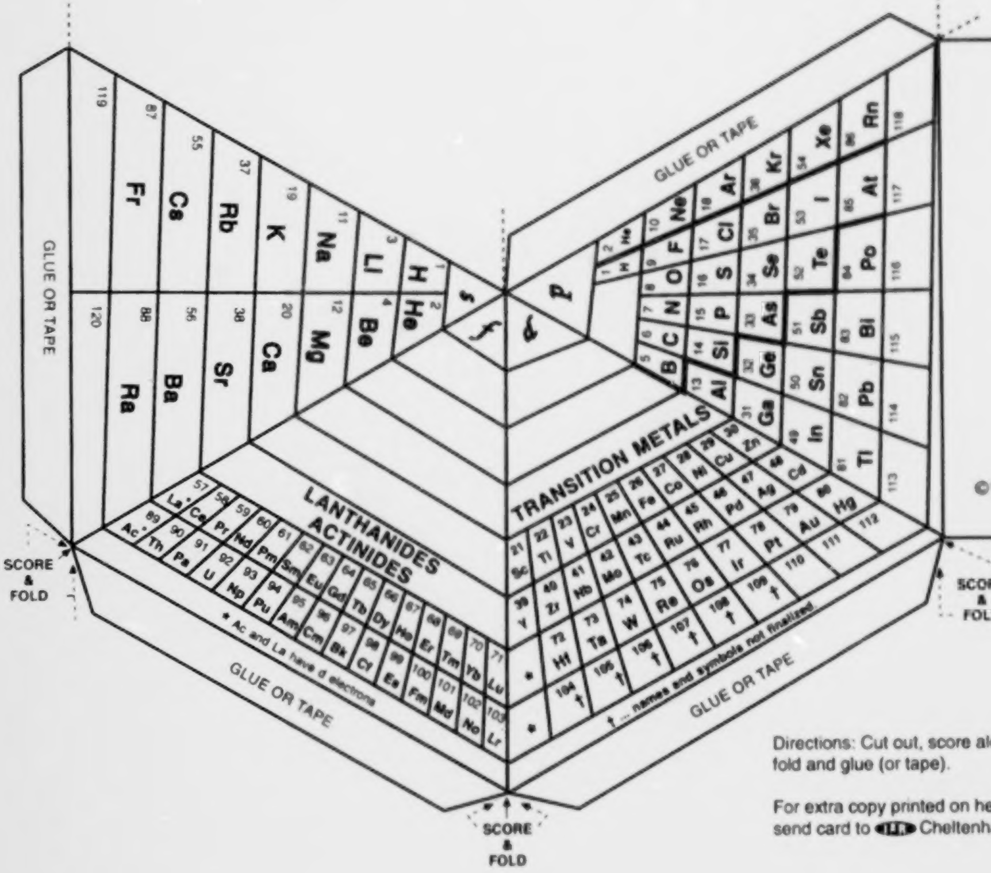
Y	T	I	V	I	T	A	G	E	N	O	R	T	C	E	L	E
L	T	B	Q	W	A	L	C	I	D	O	I	R	E	P	B	A
R	A	N	Q	Y	Y	S	C	F	I	Q	A	L	T	H	B	L
T	B	V	O	Y	I	Y	W	M	J	B	W	W	T	N	G	K
J	A	S	R	B	G	M	L	S	A	K	U	Q	J	F	P	A
Z	Q	R	R	R	L	Q	E	Q	N	C	A	D	O	E	A	L
Y	Z	N	A	B	H	E	W	E	P	O	J	T	R	D	G	I
X	L	U	D	Z	N	K	G	N	N	L	L	I	H	Q	A	M
Y	W	I	H	R	Q	O	B	A	S	B	O	S	Q	B	Z	E
L	X	V	M	J	L	R	W	R	S	D	Q	H	V	P	J	T
C	E	U	L	A	T	O	M	I	C	R	A	D	I	U	S	A
O	D	C	H	K	F	H	B	O	E	Q	E	N	N	F	N	L
I	O	N	I	Z	A	T	I	O	N	E	N	E	R	G	Y	S

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Half the distance of a single bond between two atoms of an element (2 words).
2. Measure of the attraction of an atom for a pair of shared electrons.
3. Energy required to remove the most loosely held electron of an atom in the gas phase (2 words).
4. Helium is one (2 words).
5. Astatine is one.
6. Elements in the first group (2 words).
7. Another name for group.
8. A horizontal row on the periodic chart.
9. Transition elements are in this section of the periodic chart.
10. _____ states that chemical and physical properties of the elements are periodic functions of their atomic number (2 words).

Identification of Selected Substances

Name	Formula	Color	Odor	Combustion	Mixes w/ water	Specific Test
Hydrogen	H_2	None	None	Burns	No	Burns explosively
Oxygen	O_2	None	None	Supports burning	No	Glowing splint flames
Carbon dioxide	CO_2	None	None	No	Yes	Turns lime-water milky
Nitrogen dioxide	NO_2	Brown	Choking	No	Yes	Mixed with water turns litmus red
Bromine	Br_2	Brown	Choking	No	Yes	Liquid at room temperature
Chlorine	Cl_2	Pale green	Choking	No	Yes	Forms yellow mixture with water
Ammonia	NH_3	None	Ammonia water	No	Yes	Mixed with water turns litmus blue
Sulfur dioxide	SO_2	None	Suffocating	No	Yes	Mixed with water turns litmus red
Hydrogen sulfide	H_2S	None	Rotten eggs	No	Yes	—
Hydrogen chloride	HCl	None	Choking	No	Yes	Mixed with water turns litmus red
Water vapor	H_2O	None	None	No	Yes	Cobalt chloride paper turns pink


GLUE OR TAPE
s
p
d
f
H He
Li Be
Na Mg
K Ca
Rb Sr
Cs Ba
Fr Ra
LANTHANIDES
ACTINIDES
TRANSITION METALS
* Ac and La have d electrons
† ... names and symbols not finalized
SCORE & FOLD
GLUE OR TAPE
GLUE OR TAPE
SCORE & FOLD
SCORE & FOLD
GLUE OR TAPE

THE PERIODIC PYRAMID

By Charles E. Gragg,

Burroughs Wellcome Co.
Research Triangle Park, NC

Each side represents
one type of atomic orbital
s, p, d, f
*Note exceptions

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Cheltenham, PA, for Charles E. Gragg

Directions: Cut out, score along fold lines, fold and glue (or tape).

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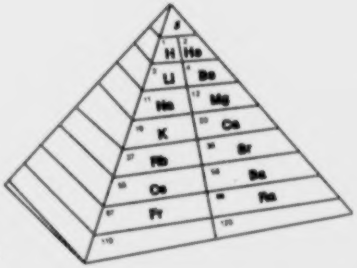
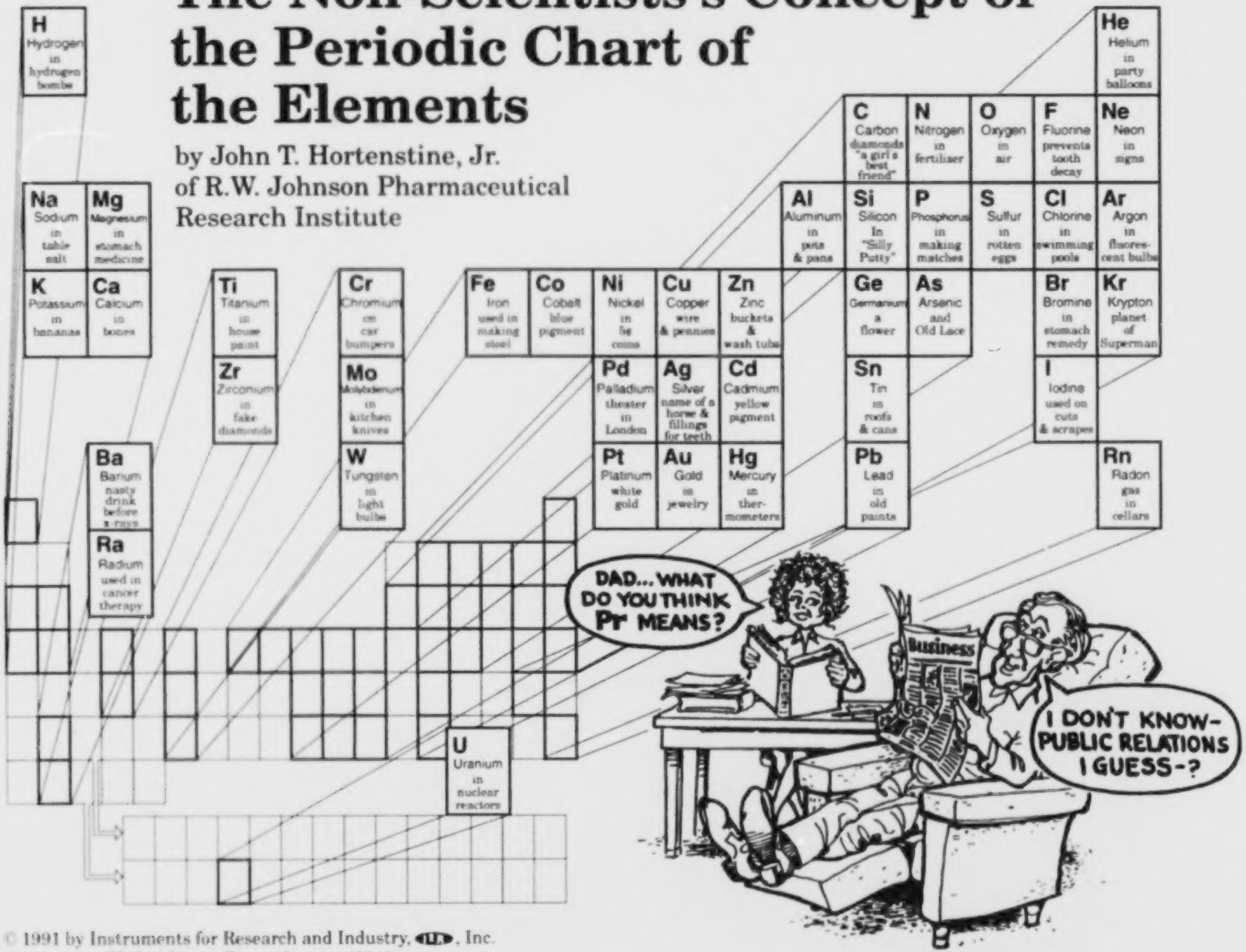

s
H He
Li Be
Na Mg
K Ca
Rb Sr
Cs Ba
Fr Ra

The Non-Scientists's Concept of the Periodic Chart of the Elements

by John T. Hortenstine, Jr.
of R.W. Johnson Pharmaceutical
Research Institute


H Hydrogen in hydrogen bombs
He Helium in party balloons
Na Sodium in table salt
Mg Magnesium in stomach medicine
K Potassium in bananas
Ca Calcium in bones
Ba Barium nasty drink before x-rays
Ra Radium used in cancer therapy
Ti Titanium in house paint
Zr Zirconium in fake diamonds
Cr Chromium on car bumpers
Mo Molybdenum in kitchen knives
W Tungsten in light bulbs
Fe Iron used in making steel
Co Cobalt blue pigment
Ni Nickel in 5¢ coins
Cu Copper wire & pennies
Zn Zinc buckets & wash tubs
Pd Palladium theater in London
Ag Silver name of a horse & fillings for teeth
Cd Cadmium yellow pigment
Pt Platinum white gold
Au Gold in jewelry
Hg Mercury in ther-mometers
C Carbon diamonds "a girl's best friend"
N Nitrogen in fertilizer
O Oxygen in air
F Fluorine prevents tooth decay
Ne Neon in signs
Al Aluminum in pots & pans
Si Silicon In "Silly Putty"
P Phosphorus in making matches
S Sulfur in rotten eggs
Cl Chlorine in swimming pools
Ar Argon in fluores-cent bulbs
Ge Germanium a flower
As Arsenic and Old Lace
Br Bromine in stomach remedy
Kr Krypton planet of Superman
Sn Tin in roofs & cans
I Iodine used on cuts & scrapes
Pb Lead in old paints
Rn Radon gas in cellars
U Uranium in nuclear reactors
DAD... WHAT DO YOU THINK Pr MEANS?
Business
I DON'T KNOW- PUBLIC RELATIONS I GUESS-?

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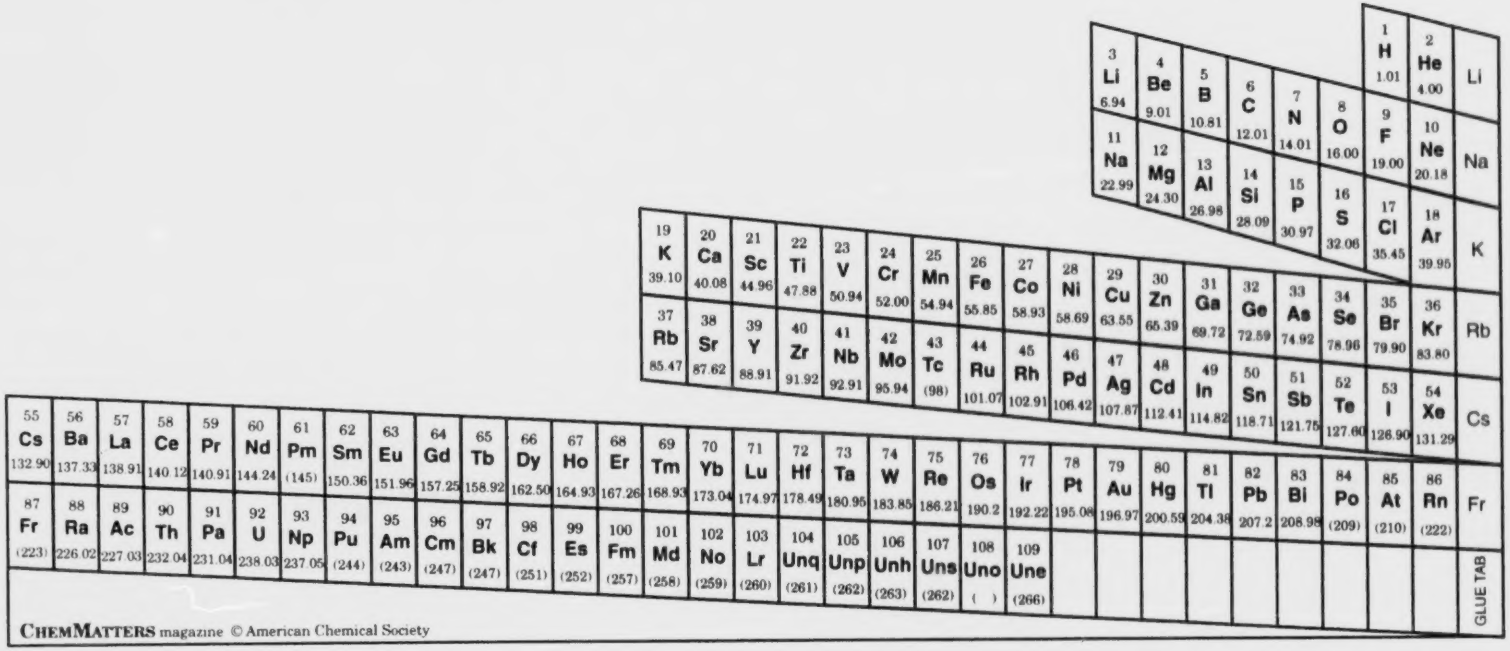
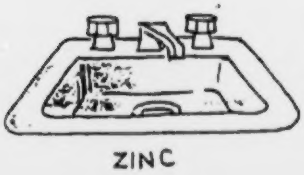
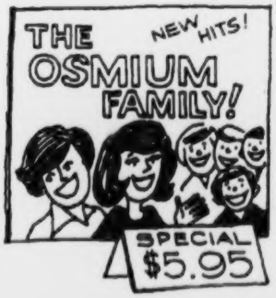
Spiral Periodic Chart Cut-Out


3 Li 6.94
4 Be 9.01
5 B 10.81
6 C 12.01
7 N 14.01
8 O 16.00
1 H 1.01
2 He 4.00
Li
11 Na 22.99
12 Mg 24.30
13 Al 26.98
14 Si 28.09
15 P 30.97
16 S 32.06
9 F 19.00
10 Ne 20.18
17 Cl 35.45
18 Ar 39.95
Na
K
19 K 39.10
20 Ca 40.08
21 Sc 44.96
22 Ti 47.88
23 V 50.94
24 Cr 52.00
25 Mn 54.94
26 Fe 55.85
27 Co 58.93
28 Ni 58.69
29 Cu 63.55
30 Zn 65.39
31 Ga 69.72
32 Ge 72.59
33 As 74.92
34 Se 78.96
35 Br 79.90
36 Kr 83.80
Rb
37 Rb 85.47
38 Sr 87.62
39 Y 88.91
40 Zr 91.92
41 Nb 92.91
42 Mo 95.94
43 Tc (98)
44 Ru 101.07
45 Rh 102.91
46 Pd 106.42
47 Ag 107.87
48 Cd 112.41
49 In 114.82
50 Sn 118.71
51 Sb 121.75
52 Te 127.60
53 I 126.90
54 Xe 131.29
Cs
55 Cs 132.90
56 Ba 137.33
57 La 138.91
58 Ce 140.12
59 Pr 140.91
60 Nd 144.24
61 Pm (145)
62 Sm 150.36
63 Eu 151.96
64 Gd 157.25
65 Tb 158.92
66 Dy 162.50
67 Ho 164.93
68 Er 167.26
69 Tm 168.93
70 Yb 173.04
71 Lu 174.97
72 Hf 178.49
73 Ta 180.95
74 W 183.85
75 Re 186.21
76 Os 190.2
77 Ir 192.22
78 Pt 195.08
79 Au 196.97
80 Hg 200.59
81 Tl 204.38
82 Pb 207.2
83 Bi 208.98
84 Po (209)
85 At (210)
86 Rn (222)
Fr
87 Fr (223)
88 Ra 226.02
89 Ac 227.03
90 Th 232.04
91 Pa 231.04
92 U 238.03
93 Np 237.05
94 Pu (244)
95 Am (243)
96 Cm (247)
97 Bk (247)
98 Cf (251)
99 Es (252)
100 Fm (257)
101 Md (258)
102 No (259)
103 Lr (260)
104 Unq (261)
105 Unp (262)
106 Unh (263)
107 Uns (262)
108 Uno ( )
109 Une (266)
GLUE TAB

CHEMMATTERS magazine © American Chemical Society

Some Very Uncommon Elements

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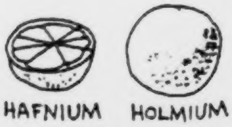
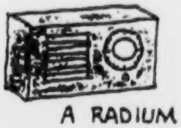
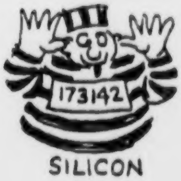

HAFNIUM
HOLMIUM

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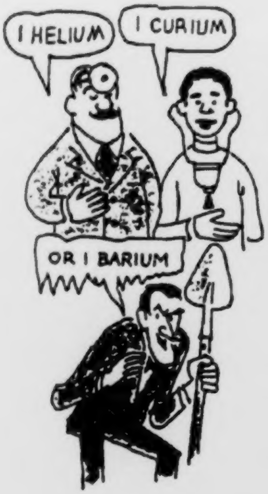
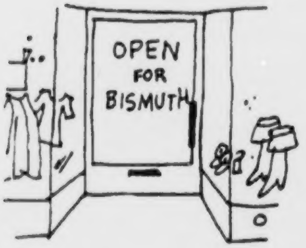
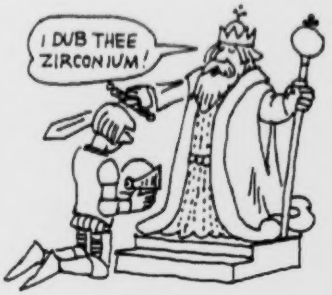
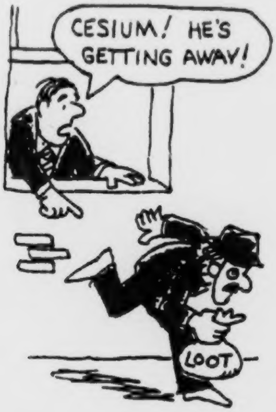

I HELIUM
I CURIUM
OR I BARIUM


OPEN FOR BISMUTH


I DUB THEE ZIRCONIUM!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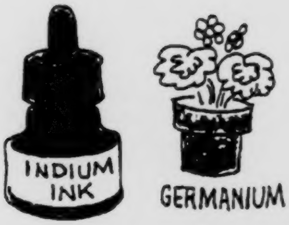
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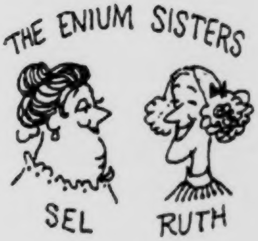
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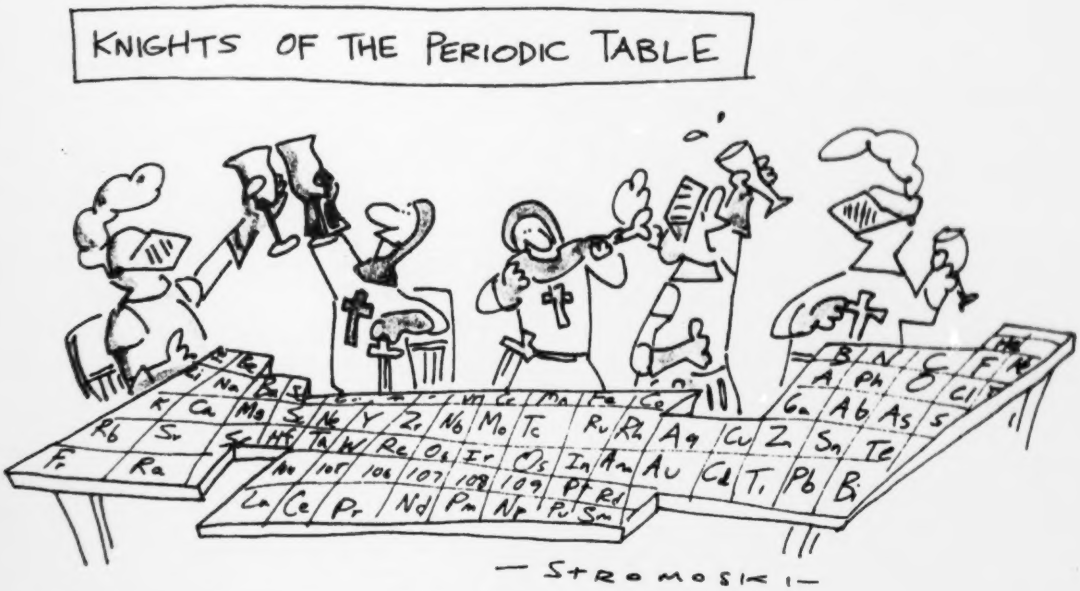
KNIGHTS OF THE PERIODIC TABLE
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— STROMOSKI —

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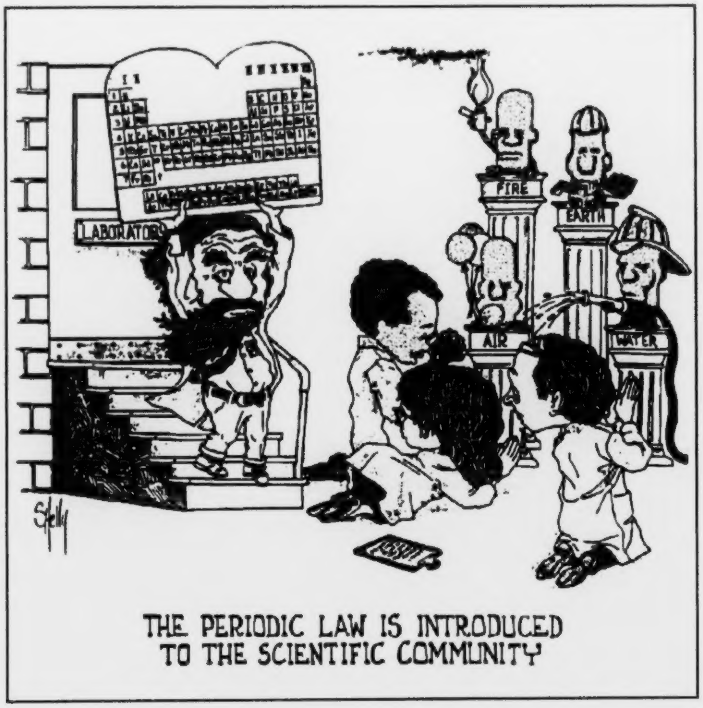
LABORATORY
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THE PERIODIC LAW IS INTRODUCED
TO THE SCIENTIFIC COMMUNITY

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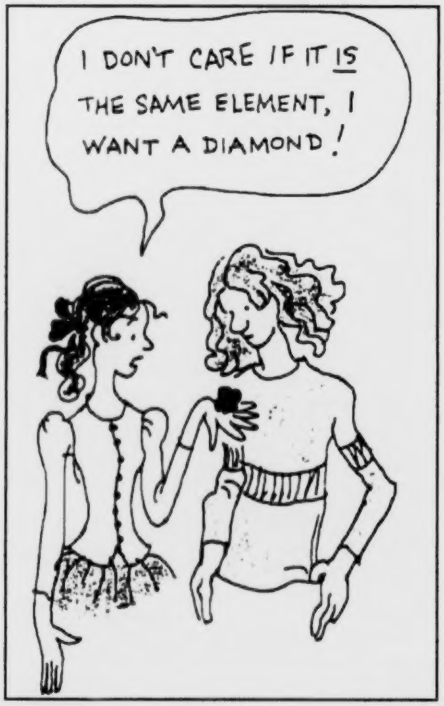
I DON'T CARE IF IT IS THE SAME ELEMENT, I WANT A DIAMOND!

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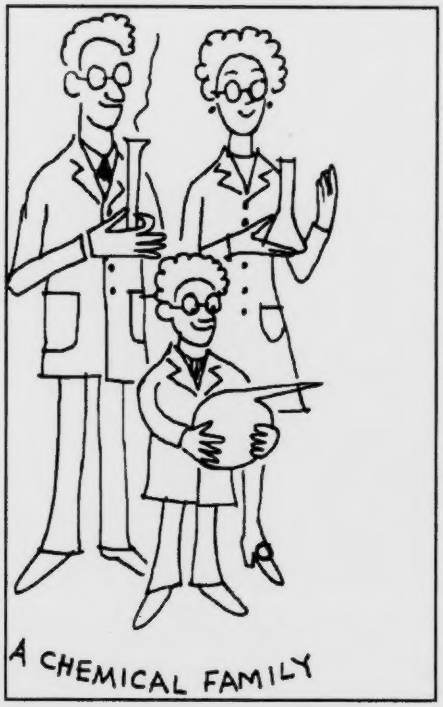
A CHEMICAL FAMILY

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A SourceBook Module

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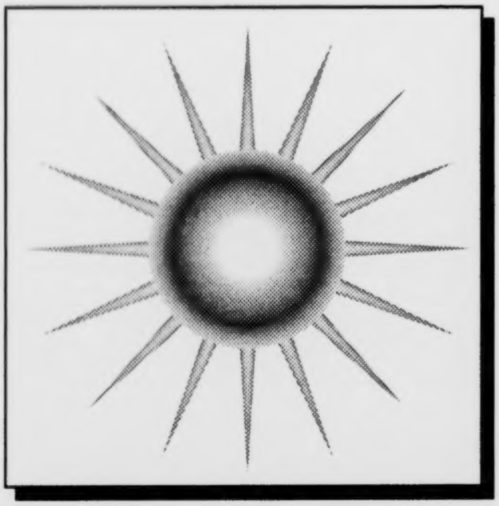
PHOTOCHEMISTRY



ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

This module emphasizes the importance of photochemical reactions and related background material necessary to understand their importance and nature. The concepts of color based on the absorption or emission of photons of specific wavelengths, of light initiating chemical reactions (sometimes unique ones) or electron flow in solid state devices, of chemical reactions producing light, of fluorescence and phosphorescence, and of quantum yields are all part of the module. Practical aspects of photochemistry include lasers, solid state solar energy devices, fluorescent lights, fabric whiteners, and new dental polymers, to name just a few. Photosynthesis is one of the most desirable photochemical reactions; ozone depletion and smog are among the least desirable. The photochemistry associated with these processes is included in this module.

PLACE IN THE CURRICULUM

This is an optional module for late in the school year, possibly after material supported by the *Reaction Rates* module is covered, although parts of the module could be used earlier as desired.

CENTRAL CONCEPTS

1. Color depends upon the selective absorption of photons (light quanta) of different wavelengths by chromophores, *i.e.*, chemical groups that absorb light. Emission of photons at specific wavelengths can also produce color. Visible light consists of photons with wavelengths between about 400-700 nm (or wavenumbers between 25,000-14,000 cm^{-1}); similar electromagnetic radiation occurs at both higher energies (<400 nm: ultraviolet, X-rays, and gamma-rays) and lower energies (>700 nm: infrared, microwave, and radiowaves).
2. Light can initiate chemical reactions by providing excited state molecules (or ions) that have sufficient energy to overcome the activation energy of the analogous thermal reactions (*e.g.*, the violent reaction that produces HCl from H_2 and Cl_2 in the presence of light). (See *Laboratory Activity 1* and *Demonstration 3*.)
3. Light activation can alter the pathways of chemical reactions, *i.e.*, provide alternate pathways to thermal reactions, different products by irradiation at different wavelengths, and the energy necessary for carrying endothermic reactions to completion.
4. Light can be used to facilitate electron flow (photoelectric effect; solar cells, see *Demonstration 4*.)
5. Light energy and the reactions activated by it can be both helpful and harmful. The protective ozone layer has been (and still is) produced photochemically from oxygen, but the layer can be destroyed by reactions of ozone with the products of chlorofluorocarbon photochemical reactions. Some polymers are synthesized through photoinitiation, while some are destroyed by photodegradation. Smog is also produced by photochemical reactions. On the other hand, the oxygen we breathe is produced by photosynthesis in plants; in fact, our entire food chain is dependent on photosynthesis (see *Biogeochemical Cycles* module).

6. Phosphorescence is effectively the emission of light (from a material exposed to radiation) that persists as an afterglow after removal of the radiation; fluorescence involves almost instantaneous (*ca.* 10^{-8} s) emission after the absorption of radiation.
7. Photoactivation in a solid crystalline material (such as AgBr) can be amplified by a factor of 10^9-10^{10} through energy transfer and is the basis of photography (see *Laboratory Activity 1*).
8. Photochemistry in the retina of the eye allows the transfer of visual images to the brain.
9. Quantum yield is a measure of photochemical reaction efficiency. A quantum yield of one means one chemical event per each photon absorbed. For chain reactions, a quantum yield greater than one is possible. Because of back reactions, fluorescence, and radiationless decay, a quantum yield much less than one is most common.
10. Lasers depend on photoactivation.
11. Some chemical reactions can produce light. This phenomenon is called chemiluminescence (see *Demonstration 1*).
12. The energy change (ΔE) associated with light absorption or emission is related to the frequency of the light absorbed or emitted (ν) by Planck's constant (h), *i.e.*, $\Delta E = h\nu$.

RELATED CONCEPTS

See *Reaction Rates* module.

RELATED SKILLS

1. The ability to read carefully and follow instructions precisely.
2. The dexterity to handle potentially hazardous solutions.
3. The only mathematical skills required for this module are arithmetic skills necessary for formulas, equations, and units.
4. Some of the concepts in this module are more easily understood if students (and teacher) have some background in energy level diagrams and/or in physics.

PERFORMANCE OBJECTIVES

After completing their study of photochemistry, students should be able to:

1. define photochemistry.
2. identify the major regions of the electromagnetic spectrum.
3. describe the basic principles of the photographic process—light activation, development, fixation.
4. explain the reason why photographic developer leaves some silver halide unchanged.
5. give examples of photochemical processes at work in everyday life.
6. describe the basis of chemiluminescent light sticks.
7. distinguish between fluorescence and phosphorescence.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: The Chemistry of Photography—A Survey

Introduction

Photography is based upon the fact that light can activate certain materials, in this case silver halides, to make them more sensitive to chemical reactions.

Purpose

To investigate the effect of light on silver salts; to investigate the effect of a developer on irradiated *vs.* nonirradiated silver salts; and to investigate the effect of sodium thiosulfate fixer on silver salts.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Be careful not to spill any of the silver nitrate solution on yourself, your clothes, your books, or the laboratory bench tops. Exposure to light will darken the silver material, and dark stains will form.
3. Any chemical spills should be cleaned up immediately.
4. Accidental spills on your body or clothes should be washed with large amounts of water and the incident reported to your teacher.
5. Dispose of the chemicals as directed by your teacher.

Procedure

Part I. The Effect of Light on Silver Salts

1. Number four strips filter paper from 1 to 4 with a pencil. Place 20 mL 0.1 M silver nitrate in a 250-mL beaker. Saturate each piece of filter paper with the silver nitrate solution, remove it with forceps, and place it on a piece of paper towelling. Cover the filter paper strips with another paper towel while obtaining the sodium halide solutions.
2. Place 10 mL each 0.1 M sodium chloride, 0.1 M sodium bromide, and 0.1 M sodium iodide in three separate 250-mL beakers. Quickly dip one of the pieces of silver nitrate saturated filter paper in the first beaker and remove it. Repeat with fresh strips of $AgNO_3$ paper for each of the other two beakers. Place all four pieces (including the piece dipped in silver nitrate only) on paper towelling. Record the appearance of each piece of filter paper. Place an opaque nonmetallic object, such as a button, on each piece of filter paper and place them in the direct rays of the sun or other bright, intense light source. Observe the appearance of the paper after 20 min. Your teacher may indicate a different time period based on the light source available to you. Save the sodium halide solutions for Parts II and III.

Part II. The Development Process

3. In a darkroom, prepare two strips of filter paper as in Steps 1 and 2 using the solution that gave you the best results in Part I. Label the strips II and III. *You do not want to expose them to general light prior to irradiation as the development process markedly increases the reactivity of light-sensitized silver halide paper.*

4. Place an opaque object in the center of each, and turn on the bright room lights for 1 min. Turn off the lights. Again, your teacher may modify this time period based on the light level in your room.
5. In the darkened room, immerse the filter papers in a 250-mL beaker containing 150 mL sulfite/glycine developing solution for 10 min. Record the changes in the papers over the 10-min period. [If working in groups, one person can do Step 7 during this time.]
6. Remove the papers from the developer solution, and place the paper labeled II on paper towelling on the table to compare with III after it has been fixed.

Part III. The Fixing Process

7. Properly dispose of the sodium halide solutions remaining in the other beakers. Wash the beakers, then thoroughly rinse them with distilled water. Place 100 mL sodium thiosulfate fixing solution in one of these 250-mL beakers. Add some (50-100 mL) distilled water to the other two beakers.
8. Rinse the paper strip labeled III in one of the beakers of distilled water. Then place the strip labeled III in the 250-mL beaker containing the fixer solution for 5 min.
9. Rinse strip in the other beaker of distilled water and place it on paper toweling.
10. Turn on the room lights and record the changes in II and III.
11. Thoroughly wash your hands before leaving the laboratory.

Data Analysis

1. What is the difference between the area covered by the opaque object and the rest of the filter paper in *Part I*?
2. What color is each of the exposed areas in *Part I*?
3. How does the silver nitrate saturated paper respond to light?
4. What sodium halide salt did you choose for *Parts II* and *III*? Why?
5. How are the papers from *Part II* and *Part III* similar to those from *Part I*? How are they different?
6. What does the final "room light" step do to the papers labeled II and III?

Implications and Applications

1. What caused the darkening of the filter papers?
2. What color is associated with each halide once it has been exposed to the light?
3. How different was the response of the silver nitrate saturated paper? Why did we observe this one?
4. What did the developer do to the silver compounds?
5. What is the main difference between the papers that were developed in *Parts II* and *III* and those that were not developed in *Part I*?
6. What does the final step of *Part II* (turning on the lights) tell you about the process? Must we complete another step to make the image permanent?
7. What causes the difference in the response of the fixed paper *vs.* the nonfixed paper to room light?
8. **Final Challenge.** What chemical reaction takes place with the silver halide in the development process?

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: The Chemistry of Photography—A Survey

Major Chemical Concept

The major chemical concept being explored here is the response of silver halides to light and the subsequent chemical reduction in the photographic process.

Level

General high school chemistry.

Expected Student Background

The major prerequisite for the laboratory is the ability to read, follow instructions, make observations, and safely handle hazardous materials. Related concepts necessary for understanding this exercise include: atomic structure, ions, formula writing, equations, and an understanding of oxidation and reduction. Students should understand what the general term "halide" means.

Time

The time needed to complete the laboratory is dependent upon the intensity of light used and whether you wait until all steps are completed in one part before going on to the next. A real timesaver is to have student groups do the various parts and share the results. A darkroom is not absolutely necessary. You can manage by just darkening the classroom and using a bright desk lamp. *Part I* takes about 40 min. *Part II* takes about 20 min, and *Part III* requires about 30 min.

Safety

Read the *Safety Considerations* in the *Student Version*. When using silver nitrate be careful not to spill any on yourself as dark stains will form in time due to the exposure of the silver material to light. You should then wash your hands with a small amount of sodium thiosulfate solution followed by plenty of soap and water.

Materials (For 24 students working in pairs)

Nonconsumables

48 Beakers, 250-mL
12 Graduated cylinders, 10-mL
12 Forceps
Strong light source
Solid, nonmetallic, opaque objects that will cast definite shadows

Consumables

0.10 M Silver nitrate, $AgNO_3$, 250 ml (4.2 g $AgNO_3$ per 250 mL solution)
0.10 M Sodium chloride, NaCl, 150 mL (0.9 g NaCl per 150 mL solution)
0.10 M Sodium bromide, NaBr, 150 mL (1.5 g NaBr per 150 mL solution)
0.10 M Sodium iodide, NaI, 150 mL (2.3 g NaI per 150 mL solution)
12 M Hydrochloric acid, HCl, 200 mL
144 Strips of filter paper (*ca.* 2 x 10 cm or cut 11 cm circles into 3 or 4 pieces)
Paper toweling
Developer solution, 400 mL (prepared by mixing 12 g glycine, 60 g sodium sulfite [Na_2SO_3], 60 g sodium carbonate [Na_2CO_3], and sufficient distilled water to make 400 mL of solution. You can use a commercial developer if you wish.)
Fixer solution, 400 mL (prepared by mixing 100 g sodium thiosulfate pentahydrate [$Na_2S_2O_3 \cdot 5H_2O$], 10 g sodium metabisulfite [$Na_2S_2O_5$], and sufficient distilled water to make 400 mL solution. Use a commercial fixer if you wish.)

Advance Preparation

The most important thing you must attend to is to do the laboratory yourself with the materials you intend to use. If you find yourself without a strong light source and a total lack of sunlight, you can use an overhead projector light. You may need to use longer time periods if your light source is not very strong. Just be sure to place the opaque object on the glass top and the filter paper on top of the opaque object; the idea is to cast a shadow on the filter paper. If the experiment extends over 3 or 4 days, fresh solutions should be prepared.

Pre-Laboratory Discussion

1. Your first task is to be sure students understand the safety instructions.
2. It is important that you demonstrate the proper placement of the opaque object between the paper and the light source.
3. It is not necessary to soak the paper too long in any one solution. Indeed you should not leave it in the halide salts for more than 30s.

Teacher-Student Interaction

At the beginning of the laboratory you should monitor the use of forceps to handle the filter paper.

Anticipated Student Results

See answers to *Data Analysis* and *Implications and Applications*.

Data Analysis

1. In Step 2 of *Part I*, the wet strips of filter paper will discolor in the light. The area covered by the opaque object will not.
2. The color of the filter paper will depend on the solution it was dipped in. The NaCl-dipped paper will become dark purple, the NaBr-dipped one will become a lighter purple, and the NaI-dipped one will become a brownish yellow.
3. Silver nitrate will turn a blush-peach color.
4. The students should choose NaCl for *Parts II* and *III* as it should provide a better contrast in *Part I*.
5. In procedure Step 5 of *Part II*, the wet strips of filter paper will show discoloration in areas where the light struck the paper and not in the area covered by the opaque object, just as in *Part I*. The use of a developer in *Parts II* and *III* makes the discoloration occur much faster. The papers become black instead of the color observed in *Part I*.
6. When the room lights are turned on in Step 10, the spots on Paper II that is light in color darkens slowly; however, the spot on Paper III does not change.

Answers to Implications and Applications

1. The substances on the filter paper were darkened by the action of light on the light-sensitive silver compounds.
2. The colors of the areas exposed to light are: silver chloride = purple, silver bromide = lighter purple-brown, silver iodide = brownish yellow.

3. The silver nitrate-saturated paper took a long time to discolor in the light and turned blush peach. Silver nitrate was used to show that the silver halides darkened and reacted faster.
4. The developer caused the reduction of the silver halide to elemental silver. Finely divided silver is black, which is the observed color.
5. The main difference between the papers from *Part I* and the other two parts is the degree of darkening and the fact that the latter two are so much blacker because the developer reduced the silver halide to elemental silver.
6. The final step (turning on the lights) caused the spot to discolor because the light exposed it while some silver halide was still on the paper. This tells us that a step to make the image permanent (or fixed) is required. The step is called "fixing."
7. The fixed paper did not lose the white spot while the nonfixed one did. This is because the fixing process removes the nondeveloped (nonlight-activated) silver halide from the paper so it cannot react and darken.

   $$AgBr(s) + 2S_2O_3^{2-}(aq) \rightarrow [Ag(S_2O_3)_2]^{3-}(aq) + Br^-(aq)$$

   The sodium metabisulfite in the fixer solution acts as a stabilizer against oxidation and bacterial growth.
8. The developer provides electrons to reduce the silver bromide:

   $$AgBr(s) + e^- \rightarrow Ag(s) + Br^-(aq)$$

   If you used the sulfite/glycine developer, the sulfite is oxidized to sulfate:

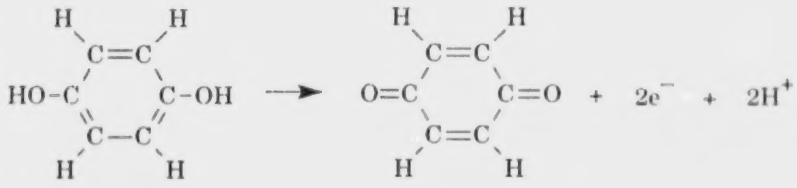
   $$SO_3^{2-}(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$$

   The net ionic equation for the total reaction is:

   $$2AgBr(s) + SO_3^{2-}(aq) + H_2O(l) \rightarrow 2Ag(s) + SO_4^{2-}(aq) + 2H^+(aq) + 2Br^-(aq)$$

   (The glycine complexes the silver of the silver halide making it more soluble and thus more easily reduced, and the carbonate neutralizes the protons formed in the sulfite oxidation.)

   When the hydroquinone commercial developer reduces the silver bromide, the hydroquinone is oxidized to quinone, which can be represented as:



Hydroquinone
$(C_6H_6O_2)$

Quinone
$(C_6H_4O_2)$

Overall:

$$2AgBr(s) + C_6H_6O_2(aq) \rightarrow 2Ag(s) + C_6H_4O_2(aq) + 2H^+(aq) + 2Br^-(aq)$$

Commercial developers contain other ingredients such as bases to neutralize the protons released in the above reaction.

Possible Extensions

Extensions are limited only by time. It is possible to relate exposure times to image intensity; temperature to development times and contrast; developer concentration to development times and contrast and more. Students can also investigate the different types of black and white film available commercially.

Assessing Laboratory Learning

Laboratory Practical. Have students expose film, process it, and make prints.

Written examination:

1. Why does the area where the opaque object was placed become dark when the object is removed from the paper that has not been fixed? Why does it stay white on the fixed paper? *[In the first case, when you remove the object the light shines on the silver halide that was previously under the object. Having been exposed to the light, the silver halide will be developed. The fixing process removes the silver halide from the paper, so in the second case, the silver halide was no longer there to react with the light; therefore, the paper stays white when exposed to light.]*
2. Write equations to represent the reactions of the developer with the silver halides. *[See the answer to Question 8 of* Implications and Applications *above.]*

DEMONSTRATIONS

CAUTION: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Chemiluminescence

Purpose

Lightsticks can be used to demonstrate chemiluminescence as well as the temperature dependence of the chemical reaction causing chemiluminescence.

Materials

3 Cyalume lightsticks (available through Flinn Scientific, Aldrich Chemical, Edmund Scientific, Markson Science, safety equipment suppliers, some hardware stores, some auto supply houses, *etc.*)
2 Beakers, 500-mL
Ice, 500 mL
Hot water, 400 mL

Safety

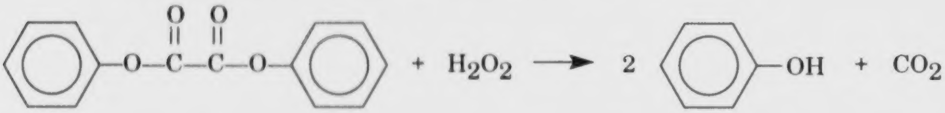
The manufacturer claims that lightstick components are nontoxic either *via* contact with skin or ingestion. In the event that the plastic case should rupture, it is advisable to thoroughly wash solutions from skin or clothing with water.

Procedure

1. Activate the chemiluminescence following the instructions on the package. Typically this requires bending the plastic tube slightly in order to break the thin-walled vial inside and then shaking the tube to ensure good mixing.
2. To demonstrate the temperature effect on the chemiluminescence reaction rate, light sticks may be heated or cooled prior to activation. Immerse one of the light sticks in an ice bath for a few minutes. Place another one into a hot water bath for a few minutes. (Because of the thick plastic lightstick container, transfer of heat is slow.) Remove the lightsticks from the baths, activate by breaking the seals, and compare the light intensities with the intensity of a third lightstick kept at room temperature. Alternatively, the lightsticks may be heated or cooled prior to luminescence demonstration.

Remarks

The chemiluminescence in the lightsticks is based on energy transfer between an intermediate formed in the reaction of phenyl oxalate ester with hydrogen peroxide (which is in the glass vial):



and an anthracene fluorescent dye molecule. The reaction intermediate is thought to be C_2O_4, an unstable molecule, which reacts with the dye molecule as follows:

$$\text{Dye} + C_2O_4 \rightarrow \text{Dye}^* + 2CO_2$$

where Dye* represents the activated dye molecule prior to emission, which can be written as:

$$\text{Dye}^* \rightarrow \text{Dye} + h\nu$$

where $h\nu$ represents the emitted photon of light of frequency ν. Further details of this reaction can be found in the "'Cool-Light' Chemiluminescence" article on which this demonstration is based (Shakhashiri *et al.*, 1981).

Flinn Scientific also sells a light stick kit that can be used as a chemiluminescence demonstration. "Chalk that Glows in the Dark" and "Chemiluminescent Glowing School Colors" demonstrations have also been published. Both involve suspected carcinogens and are not recommended for beginning high school teachers. However, light sticks are now available in several colors and intensities and may be substituted for the raw chemical demonstration of school colors. Still other demonstrations of chemiluminescence have been published in scientific journals and demonstration books, but almost all of them require chemicals that are either not readily available or are undesirable in the high school classroom.

An exception to this chemical quandary is the chemiluminescence quenching of quinine activity recently published by Sacksteder *et al.* (1990, December). Photophysics in a disco. *Journal of Chemical Education, 67*(12), 1065-1067. A demonstration version follows as *Demonstration 2.*

Demonstration 2: Quenching of Quinine Luminescence

Purpose

The quinine in tonic water is a fluorescent molecule that emits a blue glow when irradiated with a near-ultraviolet (long wavelength black light) source. The addition of salt quenches the fluorescence and acids such as lemon juice affect the fluorescence differently at different concentrations.

Materials

Ultraviolet light source
2 Erlenmeyer flasks, 500-mL
Volumetric flask, 50-mL
Graduated cylinder, 100-mL
12 Large test-tubes
Tonic water, 1 small bottle is sufficient
Sodium chloride, 1 g
Water, distilled (or tap water, if the dissolved salt level is low)

Safety

Do not stare at the black-light source. It may be injurious to your eyes.

Procedure

1. Make a "standard quinine solution" by diluting the tonic water by a factor of 10; *e.g.*, take 40 mL tonic water and dilute to 400 mL with distilled water in an Erlenmeyer flask.
2. Make a "salted quinine solution" by placing half of the standard quinine solution in a second Erlenmeyer flask containing 1 g of sodium chloride. The solution is swirled to dissolve the NaCl. (*NOTE: This procedure allows mixed solutions in Step 4 to contain the same quinine concentration.*)
3. Make a series of further dilutions of the standard quinine solution with water in large test-tubes: 12.5% (1 vol. standard : 8 vol. total solution), 25% (1 : 4), 33% (1 : 3), 50% (1 : 2), 75% (3 : 4), plus a test-tube with the full strength standard and one with just water, see following table. (These solutions can be prepared ahead of time or else the test-tubes can be marked with appropriate volume markings so that rapid dilutions are possible.) Place the quinine solution in a test-tube rack in order of ascending or descending concentration.

Dilution Table for 100 mL Solution

% Quinine	mL Standard quinine	mL H_2O
0	0	100
12.5	12.5	87.5
25	25	75
33	33	67
50	50	50
75	75	25
100	100	0

4. Make at least three sample solutions with different proportions of the standard quinine and the salted quinine solutions (*e.g.*, 9 vol. standard : 1 vol. salted, 7 : 3, and 5 : 5). Place these in a test-tube rack together with a test-tube containing the standard quinine solution and one containing the salted quinine solution.
5. In a dimly lit space (*e.g.*, in a hood or with shades drawn) irradiate the solutions, and ask students to use their eyes as fluorimeter detectors to compare the intensities of the salt-containing solutions with the dilute solutions of the standard. The intensities of the dilute solutions of the standard should be proportional to the concentrations of quinine in the solutions. Get a consensus as to the luminescence level of each of the salt-containing solutions, and come to a conclusion as to the effect of chloride ion on quinine fluorescence. *[As the chloride ion increases, the fluorescence decreases—however, it is not linear because of the nature of the mechanism. Chloride ion is said to be a quencher of quinine fluorescence.]*

6. Add a little lemon juice to the most dilute solution of the standard (12.5%). *[The intensity should increase initially, but will eventually quench the fluorescence. Adding lemon juice to the more concentrated solutions will also cause quenching. Adding of other acids will confirm that it is an acid / base effect on the quenching.]*

Remarks

The fluorescence is the result of excitation of the quinine by the ultraviolet light (*i.e.*, the energy of a photon is absorbed by the quinine)

$$\text{Quinine} + h\nu \rightarrow {}^*\text{Quinine}$$

followed by emission at a lower energy.

$${}^*\text{Quinine} \rightarrow \text{Quinine} + h\nu'$$

This allows us to see the emission in the visible region of the spectrum. [Always, fluorescence occurs at lower energy than the energy of the photons that are absorbed. Most molecules lose all of the absorbed energy as heat and show no fluorescence.]

The chloride ion acts as a quencher, *i.e.*, a quinine molecule is deactivated by a collision with a chloride ion before it has a chance to deactivate by emission of radiation.

$${}^*\text{Quinine} + \text{Cl}^- \rightarrow \text{Quinine} + {}^*\text{Cl}^-$$

$${}^*\text{Cl}^- \rightarrow \text{Cl}^- + \text{Heat}$$

The acid-base interactions are too complicated to detail here.

In practice, the amount of chloride ion in unknown solutions can be determined by quantitatively measuring the fluorescence that is obtained with known concentrations of quinine added to the solutions. An instrument called a spectrofluorimeter is used for such measurements.

More details on this demonstration can be found in the activity by Sackstedcr *et al.* (1990).

Demonstration 3: Photoreduction—The Blueprint Reaction

Purpose

To mimic the blueprint photoreaction.

Materials

0.15 M Oxalic acid, $H_2C_2O_4$, 10 mL (0.19 g of $H_2C_2O_4 \cdot 2H_2O$ per 10 mL solution)

0.10 M Iron(III) nitrate, $Fe(NO_3)_3$, 10 mL (0.40 g of $Fe(NO_3)_3 \cdot 9H_2O$ per 10 mL solution)

0.10 M Potassium ferricyanide, $K_3Fe(CN)_6$, 10 mL [0.33 g $K_3Fe(CN)_6$ per 10 mL solution]. *NOTE: $K_3Fe(CN)_6$ is commonly called potassium ferricyanide.* [Potassium hexacyanoferrate(II), $K_4Fe(CN)_6$, commonly called potassium ferrocyanide, gives a blue color immediately because of the mixed oxidation states and is not appropriate for this demonstration.]

Filter paper

Strong light source

Opaque nonmetallic solid object to obstruct the photochemical reaction

Paper toweling

Procedure

Mix the $Fe(NO_3)_3$ solution with the $H_2C_2O_4$ solution to make a solution of iron(III) oxalate. Soak a piece of filter paper in the ferric oxalate solution, and then blot the filter paper with a paper towel. Then soak the filter paper in the $K_3Fe(CN)_6$ solution. Reblot the filter paper. Place a non-metallic opaque object on the filter paper, and place it under a strong light source (or in sunlight) until a blue color is developed, which may take several minutes. Remove the filter paper from under the light source, and remove the opaque object from the paper. Rinse to remove the yellow color of the iron(III).

Whereas the above procedure has the advantage of doing the entire activity in front of students, an alternative procedure of allowing the paper to dry (in the dark or very dimly lit space) prior to the exposure to light is less messy but also less informative (see Summerlin & Ealy, 1985). If predried, a metal object, such as a key or a coin, can be used to make the pattern, whereas with moist paper, a redox reaction between the metal object and the iron(III) may occur.

Remarks

Solutions of iron(III) oxalate, $Fe_2(C_2O_4)_3$, contain several kinds of complex ions of Fe^{3+}; for example, $[Fe(C_2O_4)(H_2O)_4]^+$, which can be simplified to $Fe(C_2O_4)^+$. When irradiated, this ion can transfer an electron from oxalate to iron:

$$Fe(C_2O_4)^+ + \text{Light} \rightarrow Fe^{2+} + C_2O_4^{-\cdot}$$

The $C_2O_4^{-\cdot}$, which is a very unstable radical anion (see *Language of Chemistry*), decomposes:

$$C_2O_4^{-\cdot} \rightarrow CO_2 + CO_2^{-\cdot}$$

And a new $CO_2^{-\cdot}$ radical anion dimerizes to produce the oxalate anion:

$$2CO_2^{-\cdot} \rightarrow C_2O_4^{2-}$$

The overall reaction is:

$$2Fe(C_2O_4)^+ + \text{Light} \rightarrow 2Fe^{2+} + 2CO_2 + C_2O_4^{2-}$$

The resulting iron(II) reacts with hexacyanoferrate(III) to produce polymeric Turnbull's blue or Prussian blue:

$$Fe^{2+} + [Fe(CN)_6]^{3-} \rightarrow \{Fe[Fe(CN)_6]^-\}_n$$

where the cation that balances the polyanionic charge is largely reoxidized iron(III); *i.e.*, aqueous iron(II) is readily oxidized to iron(III). Therefore, the polymer can be represented as $\{Fe_4[Fe(CN)_6]_3 \cdot xH_2O\}_n$, where n is 14 to 16, and the preceding equation should probably be written as:

$$3Fe^{2+} + 3[Fe(CN)_6]^{3-} + Fe^{3+} + nH_2O \rightarrow \{Fe_4[Fe(CN)_6]_3 \cdot xH_2O\}_n$$

The polymeric structure consists of alternating Fe^{2+} and Fe^{3+} ions, each surrounded by six linear, bridging cyanide ions (*i.e.*, Fe–C≡N–Fe bridges) in three dimensions:

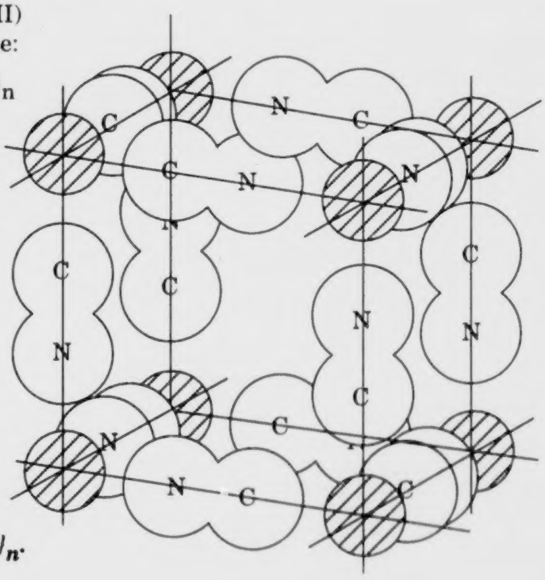


Figure 1. Polymeric structure of $\{Fe_4[Fe(CN)_6]_3 \cdot xH_2O\}_n$.

The blue color is due to electron delocalization in the polymer; that is, the iron(II) and iron(III) ions all try to become equivalent. *(TO THE TEACHER: Such intense broad absorption bands are called intervalence transfer bands.)* A similar blueprint demonstration, but using iron(III) ammonium citrate as the photoactive iron(III) component has been published (Summerlin & Ealy, 1989). These authors note that a black and white photonegative or an ordinary overhead transparency can also be used to make the pattern.

Other Demonstrations

1. **Solar cells**. A titanium dioxide solar cell can be made by heating a disc of titanium metal in a flame in order to oxidize the surface to TiO_2. Electrical leads hooked to a multimeter can be used to show that current is generated when the device is irradiated with sunlight or other bright light source. Alternatively, a standard silicon solar cell (available from Edmund Scientific) can be demonstrated in a similar manner. The current generated can be detected with a small volt/ammeter.
2. **Flashbulbs** (disposable type) can be flashed to provide an example of the generation of light by the oxidation of magnesium to magnesium oxide. *NOTE: This process is not considered chemiluminescence, but rather light generated from the heat generated in the reaction,* i.e., *incandescence.* (Magnesium ribbon burned in an open flame provides the same result.)
3. **Flame tests**, which are often performed in qualitative analysis or in the discussion of atomic energy levels, are appropriate for this module, too (see *Demonstration 1* in *Alkali Metals* and *Activity 2* in *Chemistry of Rocks, Minerals, and Gems*).
4. **Triboluminescence** demonstration has recently been suggested in *ChemMatters* (Sweeting, 1990). A Wint-O-Green Lifesaver candy disc (sugar type) is pressed between the teeth, and, in the dark, light emission can be observed. Alternatively, a pair of pliers can be used to apply the pressure (with the candy disc in a plastic bag to avoid the mess that can result if the disc shatters during the demonstration). Whereas triboluminescence produces N_2^+, which emits primarily in the ultraviolet with only a dim blue glow in the visible, increased intensity of light is obtained by energy transfer to the methyl salicylate or oil of wintergreen, which fluoresces mainly in the blue region of the spectrum. Other types of sugar-based hard candy provide weaker visible luminescence.

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. How does light initiate chemical reactions? *[Light absorption increases the energy of the molecules, placing them in excited states, which often have enough energy to overcome the activation energy for thermal chemical reactions or to cause bond breakage.]*
2. How can light alter reaction pathways? *[Light absorption can provide enough energy to access pathways that are not attainable thermally because of other lower energy reactions, or light may provide highly reactive species such as free radicals by homolytic bond breakage. Photochemical excitation often populates other than the lowest excited states. In fact, by using different specific wavelengths, different reactive pathways can be accessed.]*

3. Although ozone is decomposed photochemically it continues to exist in the atmosphere. Explain. *[When ozone is decomposed photochemically by ultraviolet light:* $O_3 + h\nu \rightarrow O + O_2$*, the O atoms can recombine with other oxygen molecules to produce more ozone. Also, oxygen molecules can absorb ultraviolet light to produce more oxygen atoms:* $O_2 + h\nu \rightarrow 2O$*, which in turn can produce more ozone:* $O + O_2 \rightarrow O_3$*.]*

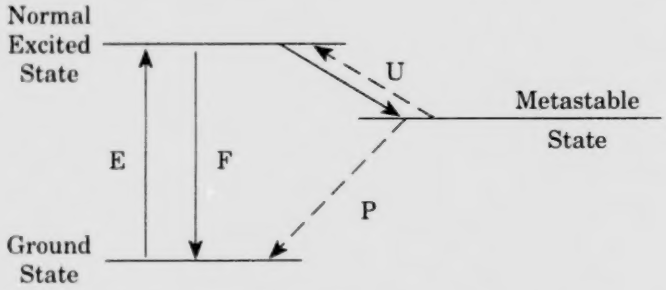
4. Both oxygen and ozone absorb in the ultraviolet region of the spectrum. Why do we need ozone to protect us from the sun's ultraviolet radiation, especially since there is so much more oxygen in the atmosphere than there is ozone? *[Oxygen absorbs primarily in the far ultraviolet region (< 200 nm) whereas ozone absorbs in the normal ultraviolet region (200–360 nm)].*

5. What processes in nature that are essential to us involve photochemistry? *[The oxygen we breathe is produced by photosynthesis; our entire food chain depends on photosynthesis; the ozone-protecting layer about the earth depends on photochemical synthesis; photochemistry in the retina of the eye allows us to see; the warmth of the earth's surface depends on photochemical absorption of light that is then dissipated as heat;* etc.*]*

6. Has modern technology found additional uses for photochemistry? *[Yes—details can be found in the* Links and Connections *section.]*

7. Why is some light emission delayed (phosphorescence)? *[In addition to normal excited states that almost immediately revert back to the ground state (fluorescence), other excited states, which are metastable, also exist. If after excitation, energy transfer to such a metastable state occurs, relaxation back to the ground state is very slow. Schematically:*

E = Excitation
F = Fluorescence (allowed, fast)
P = Phosphorescence (forbidden, slow)
U = Unfavorable (slow)



A number of other pathways also exist for the removal of the energy of excited states; therefore, most materials do not fluoresce or phosphoresce. "Allowed" and "forbidden" are quantum mechanical terms that given an indication of the probability of a reaction; a "forbidden" process is a highly improbable process.

8. How can chemiluminescence occur? *[Although the energy released in exothermic chemical reactions is typically released as heat, if the energy released has sufficient energy for light emission in the visible region of the electromagnetic spectrum, chemiluminescence is possible.]*

9. Why don't all photochemical reactions have a quantum yield of unity (one)? *[Whereas an excited state results each time a photon of light is absorbed, radiationless decay, fluorescence, and back reactions can all compete with the photochemical reaction and typically decrease the quantum yield below unity. On the other hand, a chain reaction can produce more than one product per photon absorbed, which can provide a quantum yield greater than unity.]*

Counterintuitive Examples

1. Ozone is synthesized photochemically from oxygen in the atmosphere with ultraviolet light:

$$\begin{aligned} O_2 + h\nu &\rightarrow 2O \\ O + O_2 &\rightarrow O_3 \end{aligned}$$

   However, the ozone so produced absorbs ultraviolet radiation, which leads to the reverse of the latter reaction:

$$O_3 + h\nu \rightarrow O + O_2$$

   The net effect is a low steady concentration of ozone, which continually converts ultraviolet radiation to heat; this process protects us from some of the sun's ultraviolet radiation. (The absorption maximum of oxygen is about 145 nm, whereas ozone absorbs primarily between 200 and 300 nm.) (Also, see *Societal Links and Connections* below for further complications brought about by our modern life styles.)

2. A fluorescent lamp also exhibits phosphorescence. (Turn off a fluorescent lamp in the dark, and note the phosphorescent afterglow.)

TIPS FOR THE TEACHER

Language of Chemistry

absorption of electromagnetic radiation (*e.g.*, visible light) by a chemical species is associated with the molecule going to a higher energy state.

bioluminescence chemiluminescence catalyzed by enzymes, such as that observed in fireflies, glowworms, earthworms, snails, jellyfish, angler fishes, and luminescent fungi.

chemiluminescence exothermic chemical reactions that emit light rather than heat.

chromophores chemical structural parts that absorb light. For more details of the variety of ways in which color is generated (see *Appendix*).

emission of electromagnetic radiation corresponds to a chemical species in an excited electronic state going to a state of lower energy.

fluorescence and **phosphorescence** (see *Key Question* 7).

free radical uncharged atom or a molecular fragment with an unpaired electron. Many photochemical reactions produce free radicals. Radical anions (or cations) are unstable anions (or cations) with an unpaired electron analogous to neutral free radicals. –. and +. are the symbols for singly charged radical anions and cations, respectively. The dot represents the unpaired electron.

infrared spectral region on the lower energy (longer wavelength) side of the visible region of the spectrum. This region is associated with molecular vibrations and heat.

laser (Light amplification by stimulated emission of radiation) device producing sharp intense monochromatic beams of light. Atoms or molecules in excited states can be stimulated to emit radiation that is in phase with the wave that stimulated it. This emission is an amplification of the initial wave. In a large aggregation of atoms or molecules, appropriate design allows for continued amplification to produce an intense in-phase beam of monochromatic light (see *Links and Connections–Community*).

luminescence emission of light from relatively cool systems (as opposed to incandescence, which is the emission of light from systems hot enough to glow from the excitation obtained from the heat alone). Luminescent pigments exhibit colors through emission as well as absorption, whereas other pigments exhibit color solely on the basis of absorption. (The emission occurs at lower energies than the absorption.)

phosphors materials that exhibit phosphorescence or fluorescence. Originally the term applied to microcrystalline substances, but is now applied to noncrystalline materials as well.

quantum yield ratio of the number of product molecules formed to the number of photons absorbed.

triboluminescence mechanically induced emission observed when certain crystals (*e.g.*, sugar cube, hard candy or zinc sulfide) are crushed or when a ruby or sapphire crystal is cut with a diamond saw. The blue light results from the strong electric fields that result during the formation of new crystal faces. The electrons produce N_2^+, which provides a blue luminescence. Triboluminescence also occurs when some adhesives are rapidly removed from surfaces, which also produces luminescent N_2^+.

ultraviolet spectral region on the higher energy (shorter wavelength) side of the visible region of the spectrum. This region is associated with electron energy state changes with energy transfer to molecular vibrational modes and bond breakage.

Pattern Recognition

Relate colors with wavelengths and wavelengths with energy (see *Atomic Structure* module and Figure 2). In nanometers: 700, deep red; 650, red; 600, orange; 580, yellow; 550, green; 500, green; 450, blue; 400, violet (Nassau, 1983, p. 378).

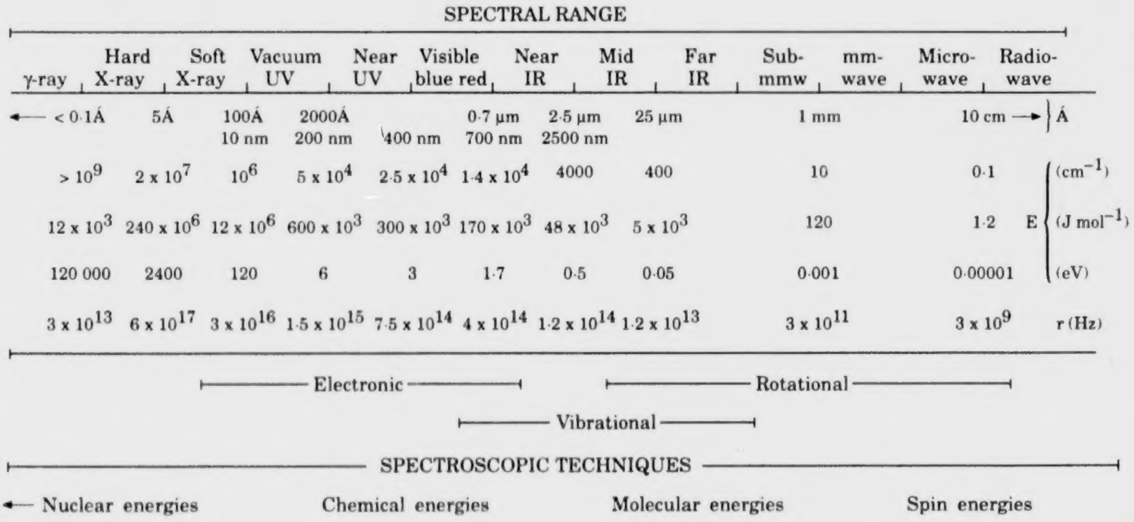

SPECTRAL RANGE
γ-ray
Hard X-ray
Soft X-ray
Vacuum UV
Near UV
Visible blue red
Near IR
Mid IR
Far IR
Sub-mmw
mm-wave
Micro-wave
Radio-wave
< 0·1Å
5Å
100Å 10 nm
2000Å 200 nm
400 nm
0·7 µm 700 nm
2·5 µm 2500 nm
25 µm
1 mm
10 cm
Å
> 10^9
2 x 10^7
10^6
5 x 10^4
2·5 x 10^4
1·4 x 10^4
4000
400
10
0·1
(cm^-1)
12 x 10^3
240 x 10^6
12 x 10^6
600 x 10^3
300 x 10^3
170 x 10^3
48 x 10^3
5 x 10^3
120
1·2
E
(J mol^-1)
120 000
2400
120
6
3
1·7
0·5
0·05
0·001
0·00001
(eV)
3 x 10^13
6 x 10^17
3 x 10^16
1·5 x 10^15
7·5 x 10^14
4 x 10^14
1·2 x 10^14
1·2 x 10^13
3 x 10^11
3 x 10^9
r (Hz)
Electronic
Rotational
Vibrational
SPECTROSCOPIC TECHNIQUES
Nuclear energies
Chemical energies
Molecular energies
Spin energies

Figure 2. Electromagnetic spectrum.

Common Student Misconceptions

1. **"The glow of a 'firefly' is from a fire."**

   The glow of a 'firefly' is a chemiluminescent reaction.

HISTORY: ON THE HUMAN SIDE

Qualitative aspects of photochemistry were known long before the development of chemistry as a science. For example, the history of chemical luminescence goes back to 1603 when Vincenzo Cascariolo, an alchemist in Bologna, Italy, heated barite (barium sulfate) and coal together and observed that upon cooling the powder exhibited a bluish glow in the dark and that the glow could be reactivated by exposure of the powder to sunlight. He apparently had made barium sulfide, though the powder became known as *lapis solaris* (sunstone) or *phosphorus* (light bearer). Now we reserve the name phosphorus for the element, which does show phosphorescence, and call microcrystalline luminescent materials **phosphors**.

Robert Boyle, the Irish chemist better known for a gas law, is given credit for first having scientifically investigated bioluminescence in 1672, when he established that the luminescence emissions from rotting wood, rotting flesh, and fireflies require air, are cold-light systems, and can be inhibited by chemical reagents such as alcohols, ammonia, and hydrochloric acid.

Photochemistry *per se* is of more recent origin (with the possible exception of the first item in the *Humor* section). Although some alchemists compared their results in sunlight *vs.* in the dark, the fact that silver halide blackening is caused by light is thought to have been discovered in the early 1700s, and the well studied photochemical gaseous reaction between hydrogen and chlorine to produce hydrogen chloride was first noted in about 1801. However, quantitative photochemistry did not develop until after Einstein's photoelectric effect pronouncements at the beginning of the 20th century provided a foundation on which others have developed an understanding of photochemical reactions.

The history of the photochemistry of photography is normally traced to Thomas Wedgwood, whose father started the famed line of pottery that bears their family name. In 1802, Wedgwood reported that paper or leather sensitized with silver nitrate provided temporary images when exposed to light using a primitive form of a camera, which was used for tracing images at that time. In France in 1826, Joseph Nicephore Niépce made a "heliograph" using the change of solubility of a special asphalt when exposed to a sunlight image over several hours. He joined Louis Daguerre in a partnership in 1829, but died in 1833. In 1837, Daguerre produced his first "daguerreotype" using a silvered copper plate exposed to iodine vapor prior to its exposure to light. After the exposure to light, mercury vapors produced a silver amalgam where the silver iodide had been reduced, and sodium thiosulfate was added to remove the unexposed silver iodide, *i.e.*, to fix the image. He and Niépce's son sold full rights for their processes to the French government in August 1839 in return for life annuities.

Meanwhile, back in England, William Fox Talbot invented a photographic process because he was unable to draw landscapes well. He was apparently unaware of the French work, or of that of Wedgwood, when in 1833-1835 he developed a light sensitive paper by soaking the paper alternately in solutions of sodium chloride and silver nitrate. He exposed the paper through a camera until the silver chloride became dark where irradiated. The unexposed portions were made less sensitive by subsequent washing with a strong sodium chloride solution. (Some of the silver chloride forms $AgCl_2^-$ and is removed.) He later switched to sodium thiosulfate at the recommendation of Sir John Herschel, another British scientist. In January 1839, when word of the French work reached England, but several months before the details were published, Talbot rushed publication of his "photogenic drawing" process and discussed it before the Royal Society. By 1840, Talbot had learned that much shorter exposures of light caused a latent image that could be developed into a real image, thus laying the foundation of the modern photographic process. The many more recent developments of photography which led to the field as we know it today are beyond the scope of this module.

HUMOR: ON THE FUN SIDE

1. Photochemists claim that their field is the oldest scientific field in the universe, and claim that God was the first photochemist when God said "Let there be light" and there was light. Photosynthesis was thereby initiated.
2. And did you hear about the little CHEMoron who thought that LASERS were people who sit around all day doing nothing?
3. See Sydney Harris' new book *Einstein Simplified: Cartoons on Science* for more humor related to this module (Harris, 1989). For example: Early scientific fraud—young Thomas Edison tried to pass off a container filled with fireflies as an incandescent bulb.
4. Message on bumper stickers:
   a. I'm in an excited state!
   b. Be a firefly: Do chemistry by night.
5. Word Search (see *Appendix* for master copy)

~~T R I B O L U M I N E S C E N C E~~
C J U P P J U M M U L W Y C R O W
X E N Q H ~~R O H P S O H P~~ M H X A
O L E D K Q B M J K A U W Y J K V
~~E C N E C S E N I M U L I M E H C~~
R K G Q E B G D W Q G Y V L A P K
E C Q V T A E J V G J N Q E D D M
V I B U W ~~R E S A L~~ W R A F C J M
L K Q W A ~~C H R O M O P H O R E S~~
I F Y R A N S Q I G O C H D V X C
S H F D R F ~~T E L O I V A R T L U~~
P N J D O M L U G L R H Q D K E X
I D Y W I Z D E M M F ~~K C N A L P~~

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Spectral region associated with molecular vibrations and heat.
2. Acronym for light amplification by stimulated emission of radiation.
3. Material that exhibits phosphorescence.
4. Kind of yield that measures photochemical reaction efficiency.
5. Emission of light through chemical reaction.
6. Mechanically induced emission of light.
7. Another name for "black light."
8. Chemical groups that absorb light.
9. Halides of this metal are light sensitive.
10. ΔE = hν is called the Einstein-_______ relationship.

Answers: 1. INFRARED 2. LASER 3. PHOSPHOR 4. QUANTUM 5. CHEMILUMINESCENCE 6. TRIBOLUMINESCENCE 7. ULTRAVIOLET 8. CHROMOPHORES 9. SILVER 10. PLANCK

6. See cartoons at the end of module.

MEDIA

1. World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.
   - a. "Number 17: The Precious Envelope," on the atmosphere, includes a discussion of the ozone layer. The videotape also has a short discussion of CO_2, photosynthesis and the greenhouse effect.
   - b. "Number 10: Signals from Within," focuses primarily on molecular spectroscopy. However, it has a good discussion of suntanning and sunscreens as well as the energy and types of electromagnetic radiation and the absorption of specific energies of radiation by molecules.
   - c. "Number 2: Color," emphasizes dyes and notes the use of fluorescent dyes in greenhouse windows to improve the quantity of energy absorbed by plants.
2. Videodiscs published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
   - a. "Depletion of Ozone," a chapter on *The World of Chemistry: Selected Demonstrations and Animations*: Disc II (double sided, 60 min.), Special Issue 4.
   - b. "Energy States and a Sound Wave Analogy" and "Energy States and Chlorophyll," two chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc I (double sided, 60 min.), Special Issue 3.
   - c. *Demonstrations in Organic Chemistry* (double sided, 60 min.). Special Issue 6.
3. Written materials available from ICE–Institute for Chemical Education, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-3033 (voice) or (608) 262-0381 (FAX).
   - a. Topics in Chemistry: *Ozone–Chlorofluorocarbons and the Hole in the Ozone Layer.*
   - b. Topics in Chemistry: *Black and White Photography.*

INSTRUMENTATION AND EQUIPMENT

1. Light sources include the sun, sunlamp bulbs, overhead projection bulbs, ultraviolet lamps, and lasers, if available for demonstration (although not necessary).
2. Solar cell and meter to show output are available from a number of supply houses, such as Edmund Scientific.

Links/Connections

WITHIN CHEMISTRY

This topic is closely associated with chemical kinetics (see *Reaction Rates* module).

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

Photochemistry is today quite closely allied with environmental science because of the ozone layer and smog problems, with physics (photophysics) in terms of the theoretical aspects of the subject, with medicine from the standpoint of links between skin cancer and irradiation, and with materials science in terms of the development of new lithographic procedures and new photoelectric and solar devices.

TO THE CONTEMPORARY WORLD

Personal

1. **Vision**. In the retina at the back of the eye, color is detected by three sets of cones at moderate to high light intensity. The cones have maximum sensitivity to different colors as shown in Figure 3, where "B," "G," and "R" are the blue, green, and red sensitive cones. At low light levels, light is detected by rods, which do not distinguish color, but do have much higher sensitivity as shown in Figure 3. The photochemical reaction thought to predominate at higher intensity levels is shown schematically in Figure 4, where rhodopsin is a condensation product between opsin, a large molecule of unknown structure, and 11-*cis*-retinal. Absorption of a photon into the so-called π^* level of the retinal allows rotation at the 11-position to the *trans* form shown on the right of Figure 4, providing a nerve impulse to the brain. A second photon can reverse the process, *etc.* The actual process involves at least two intermediates, and the process for dim light is even more complicated; however, the *cis-trans* conversion is involved in all cases. The cones are thought to have at least three different opsins, which provide the different energy absorption profiles necessary for color detection (see also *Organic Chemistry* and *Chemistry in Medicine* modules). Recently, workers at Brandeis University and Stanford University succeeded in isolating the pigments of human color vision by chemically synthesizing the genes for the pigments and expressing the genes in tissue culture. The opsin (protein) portion of these pigments is covalently linked to the chromophore, 11-*cis*-retinal through a lysine (an amino acid) side chain. Substitution of different amino acids at critical positions on the opsin chain can produce different absorption maxima and hence, different perceived colors. For example, substitution of an alanine for a serine at position 180 on the opsin results in a subtle but detectable difference in color vision tests. The consequences? Everyone does not see the same color! (See Borman, 1992, in *References.*)

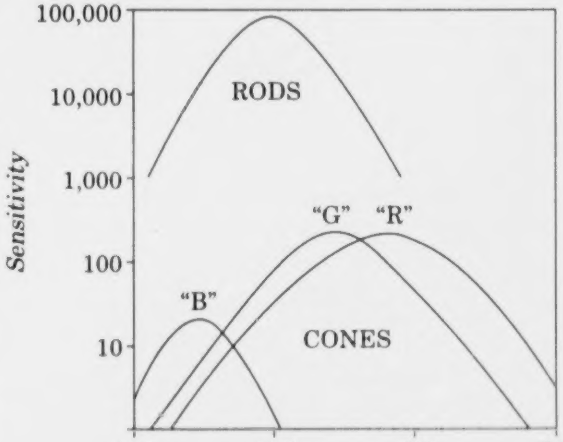


Figure 3. Sensitivity of cones and rods of the human eye.

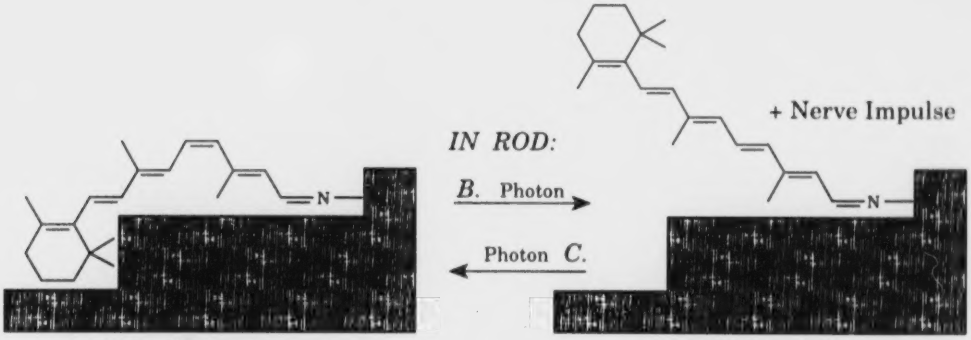


Figure 4. Light-initiated retinal chemical transformation leading to a nerve impulse.

2. **Suntanning** (and skin cancer). Exposure to the sun (or at least to ultraviolet radiation) is essential for the photochemical production of Vitamin D and for suntanning, which is based on the photochemical formation of melanin, the same pigment responsible for dark skin and black hair. However, too much ultraviolet exposure can cause skin cancer. Ultraviolet light has enough energy to modify the DNA of skin cells, which can lead to mutations, some of which are carcinogenic. *ChemCom* (Lippincott *et al.*, 1988, pp. 431–434) contains more details (see also Robbins, 1984).
3. **Photography** is probably the most common example of a photochemical process. What is not commonly understood is the chemistry of the entire process. The image is created by light exciting the electrons of silver halides. The more light a given area receives, the more electrons in that area are excited. Eventually a grain of silver bromide receives enough light to become activated and available for development. [The excited silver bromide grains can be reduced by the developer more easily than non-excited atoms.] When developed, the grains that have received enough light to be activated are reduced to black finely divided silver metal, while those areas that are not activated are not reduced. Instead, these grains of silver bromide are removed during the fixation process. In regions where only a fraction of the silver bromide grains receive enough light to be activated, gray shades develop. This light/dark contrast produces the detail in the photographic image. Because this process is based on the difference in the ability of the silver to be reduced, development times and temperatures are critical in the production of fine photographs. Excessive time spent in the developer will result in too many non-excited grains being reduced and a poor image. The second step in the process involves the fixer. The purpose of the fixer is to remove the remaining, undeveloped, silver halide compounds so the image will be "fixed" or stable (see *Laboratory Activity 1*).
4. Some **detergents** contain whiteners to make laundry appear "bright." These compounds cause the fabric to fluoresce, that is, to emit visible radiation in response to ultraviolet radiation. The "brightener" actually emits light due to absorption of ultraviolet light from the light source. A good example of this is the "blacklight" (ultraviolet source) response of white cloth or fluorescent posters.

5. **Fluorescent lamps** are glass tubes that are coated on the inside with one or more phosphors, *e.g.*, ZnS, and filled with a small amount of mercury vapor diluted with argon. Electricity passed through the gas excites the mercury atoms, which emit light, primarily in the ultraviolet and the blue end of the visible spectrum. The phosphors convert the radiation to lower energy radiation in the rest of the visible spectrum. With appropriate phosphor mixtures, daylight can be approximated. A particular advantage of the fluorescent lamp is its greater efficiency relative to the standard incandescent tungsten filament lamp, *i.e.*, a fluorescent lamp requires only about 30% as much electricity to produce the same amount of light. Fluorescent lighting is also cooler, as less energy is lost as heat. *CAUTION: Because of the mercury and of certain phosphors used in the lamps, proper disposal of the bulbs is important. Broken bulbs cause mercury contamination.*

6. When you visit the dentist and require a filling near the front of your mouth, you may have a **polymer substance** used instead of the more common silver amalgam. The reason for this is that it is possible to color the polymer to match your teeth and thus be nondetectable. These dental polymers are hardened or "dried" by shining a strong light on them. The light provides the energy necessary to initiate the polymerization reaction.

7. **Light activation** is also used to produce the masks for the elaborate silkscreen designs used on T-shirts. The silkscreen process works by having ink pushed through a fabric "screen" that has been covered in areas where no dye is wanted. To achieve the elaborate designs, a film is used. A copy of the design is placed on the film, and then it is exposed to a very strong ultraviolet light for a specific, short amount of time. This process sensitizes the film in a fashion similar to the photographic process. The film is then washed with water. The film material swells and the areas where no light struck the film are washed away. The "mask" is then placed on the screen, allowed to dry, and used.

8. **Photoelectric effect applications**. Many materials will eject electrons when irradiated with light with appropriate energies. This phenomenon is termed the photoelectric effect or photoemission. The effective wavelengths of radiation that cause this emission of electrons vary with the materials being irradiated. It is the basis of solar cells used in some of the hand-held calculators. Silicon based cells are the most commonly used solar cells. Light energy is converted to electrical energy by the silicon cells. The energy of the photon is sufficient to move an electron into a conducting band.

   Television cameras are also based on light induced electron emission at specific points on a two-dimensional solid-state photoelectric array. In color cameras, sensitivity at three different energies is required to differentiate between red, green, and blue light at each point on the array. Electronic circuitry allows for the rapid scanning of the various points of the photoelectric array.

9. **Photolithography applications**. The printed circuits used in calculators and personal computers take advantage of photochemical reactions during their manufacture. Typically, a semiconducting silicon wafer, which is the heart of the memory chip, has its surface oxidized to silica, SiO_2, which is an insulator. In order to conduct a current to the silicon, the silica must be

selectively etched. Typically, a photosensitive polymer (a resist) is placed on the silica surface and selectively irradiated through a mask. If a positive resist has been used, the irradiated portions of the resist are broken into short segments (undergo scission) and are more readily dissolved by a developer that removes the irradiated portions. Conversely, if a negative resist has been used, the irradiated portions of the resist become crosslinked and insoluble. Therefore, a developer removes the nonirradiated portions of the resist. After development, an etchant (*e.g.*, HF) etches the silica from the silicon surface in the regions in which the resist had been removed. Doping of the silicon or adding a metal for connections is thus made very specific. The present devices have many more steps and include multilayers, but photolithography is a central part of all such processes (see Figure 5, which details this procedure).

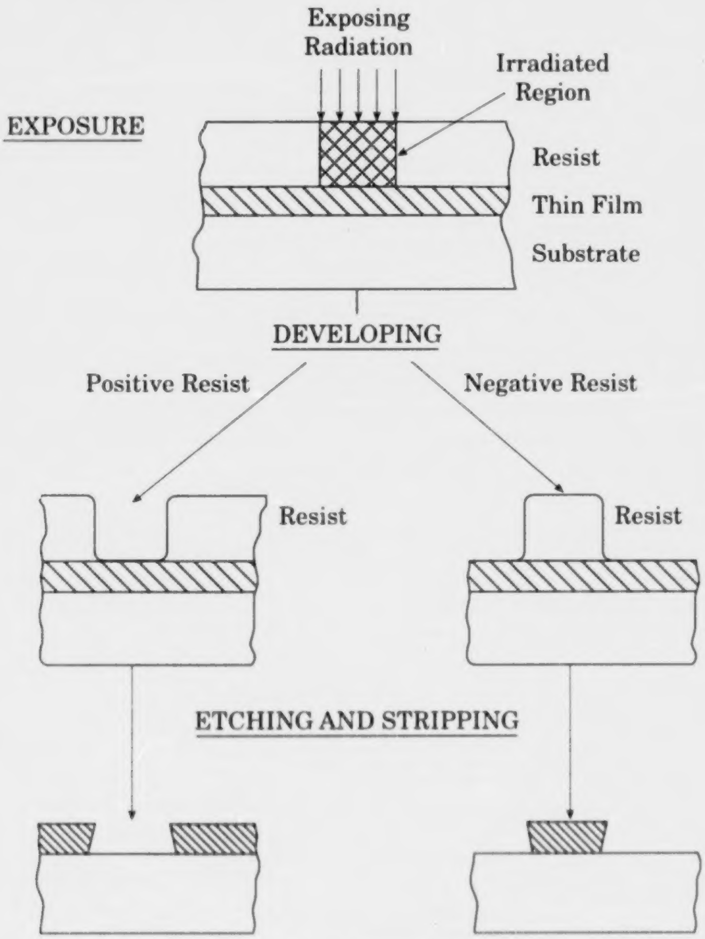


Figure 5. The scematic diagram showing the formation of a polymeric relief image using lithography.

Community

1. **Lasers**. **L**ight **a**mplification by **s**timulated **e**mission of **r**adiation devices produce sharp intense monochromatic beams of light, which are being used in a variety of devices today. Atoms or molecules in excited states can be stimulated to emit radiation that is in phase with the wave that stimulated it. This emission is an amplification of the initial wave. In a large aggregation of atoms or molecules, appropriate design allows for continued amplification to produce an intense in-phase beam of monochromatic light.

   Types of lasers in use today include the now classic photochemically activated (optically pumped) solid lasers, which have been used for precise control as in laser eye surgery, extremely high temperature melting and welding, accurate alignment measurements, holography, *etc.* A figure of such a laser is shown in Figure 6 (The New Encyclopædia Britannica, 1986).

   The heart of a typical laser consists of a rod of a material containing an ion that absorbs and emits photons (light quanta). Examples: the ruby laser—chromium(III) in an artificial sapphire crystal (Al_2O_3), and the neodymium laser—neodymium(III) in an artificial sapphire crystal. The rod has polished mirror-coated ends to reflect the laser light and has a small hole in one mirror to allow light through to the Q-switch, which reflects the light back through the crystal until it is at a high level and then opens to emit the optically pumped beam. The flash lamp that pumps the laser crystal is a high intensity lamp, either wound around the crystal as shown or positioned beside the crystal and/or focused onto the crystal. The emerging beam is coherent (monochromatic in-phase) radiation. Such laser beams are used by scientists, engineers, surveyors, and others who need sharp light beams.

Semiconductor diode lasers produce a laser beam at the junction between two dissimilar (n and p) semiconductors when a large electrical current is passed through the device. Although the power is lower than the lasers noted above, the low cost of these small efficient lasers has revolutionized telephone communications (laser light transmission though glass fibers), the record industry (audio compact disc systems), publishing (laser printers), optical memory, *etc.* (see *Materials Science* module for a discussion of semiconductors). Related types of lasers include liquid lasers, tunable dye lasers, and gas lasers. Most encyclopedias provide details on these types of lasers.

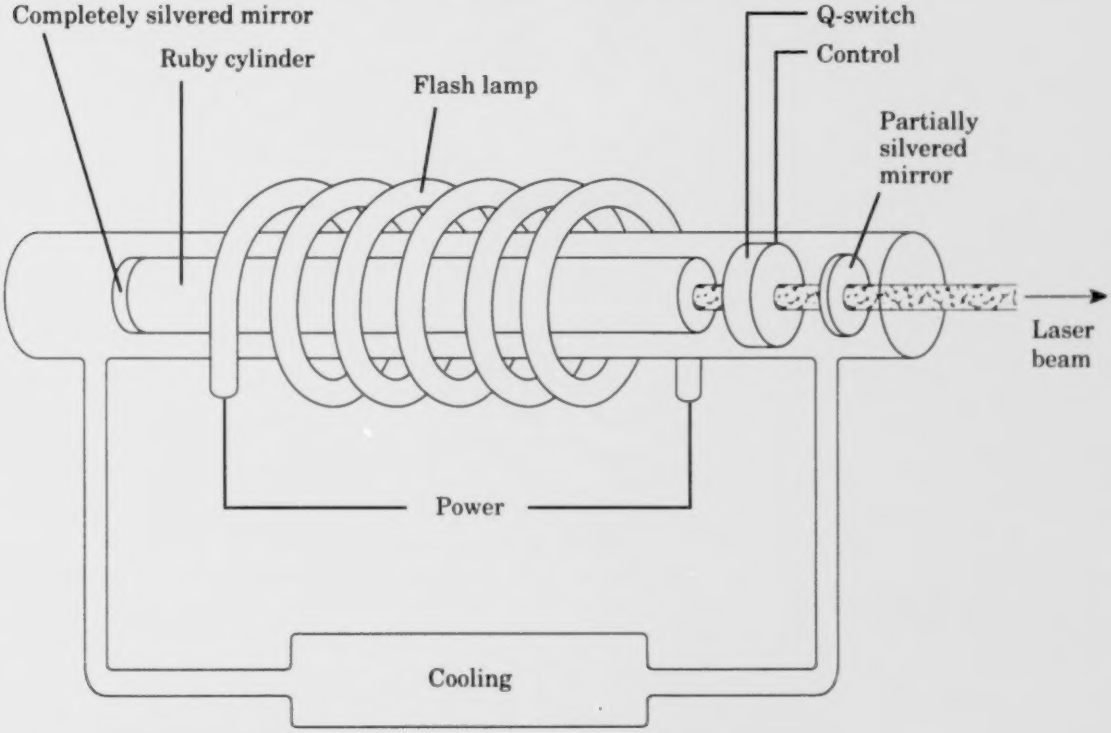


Figure 6. Q-switch, a special switching device that produces giant output pulse. In this example, a ruby laser is being used.

2. **Photochemical separation of isotopes** (uranium enrichment). The low temperature vibration of uranium-235 hexafluoride at 629 cm^{-1} can be selectively irradiated relative to the vibration of uranium-238 hexafluoride at 628 cm^{-1} or vice versa. Because of the narrow line widths possible with lasers, selective ionization followed by chemical reaction is possible. Details of this and related isotope separations can be found elsewhere (Coyle *et al.*, 1982; Andrews, 1986; Spindel & Ishida, 1991).

3. **Photoelectric effect applications** (see also *Personal* section). Xerography or photocopying is based on the localized charges that can be induced on a selenium drum by light absorption. An encapsulated graphite powder is selectively transferred to paper based on the charge differences.
4. **Photomultipliers** are electron emitters that have a set of secondary emitters built into the same device to provide amplification of the light signal and are used for light intensity measurements in a wide variety of applications, such as electric eyes, spectrophotometry, scintillation counters, facsimile transmission, and industrial controls.
5. **Photolithography** including photoengraving, photoetching, printed circuits, and other related procedures all depend on photochemical reactions. The blueprint process (*Demonstration 2*) is just one example. Photochemically reduced dichromate in the presence of a number of substances, such as egg albumin, forms insoluble crosslinked materials. The irradiated portion is the background part of the master copy, and the nonirradiated portion is the print area. The more soluble nonirradiated portion of the albumin coating is removed. The chromium polymer protects the irradiated portion of the plate when an etchant is then applied. The result is a bas-relief (gravure or intaglio) plate. Ink is applied over the entire plate and scraped from the nonbas-relief areas. Pressure and capillary action are then used to transfer the ink to the paper.

Societal (Science/Technology/Society; Current Events)

1. Smog. The automobile produces NO, which in turn is oxidized to NO_2 in the atmosphere, producing the brown color of smog. Sunlight photochemically dissociates NO_2 to NO and oxygen atoms:

   $$NO_2 + h\nu \rightarrow NO + O$$

   These oxygen atoms react with oxygen molecules to produce ozone:

   $$O_2 + O \rightarrow O_3$$

   This is the photochemical component of smog.

2. Ozone. Inasmuch as ultraviolet radiation has enough energy to break chemical bonds, life on earth is blessed with an atmosphere that absorbs much of the ultraviolet radiation. Oxygen absorbs the far ultraviolet radiation, and ozone absorbs radiation in the near ultraviolet region. However, other materials in the atmosphere are causing a decrease in the stratospheric ozone layer. Free radicals from photochemical reactions of chlorofluorocarbons (Freons), nitrogen oxides, and methane in the atmosphere can react with ozone to form oxygen gas and other oxides. To illustrate:

   $$\begin{aligned} CCl_3F + h\nu &\rightarrow Cl + \textit{etc.} \\ Cl + O_3 &\rightarrow ClO + O_2 \\ ClO + O &\rightarrow Cl + O_2 \end{aligned}$$

   Note that at the end of the sequence a Cl radical is still available to react with another ozone molecule. Estimates that 1000 ozone molecules are destroyed by each chlorine atom have been made by considering other species with which the chlorine atoms can react and their concentrations in the atmosphere (Coyle *et al.*, 1982; Lippincott *et al.*, 1988, pp. 370–372).

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Appendix

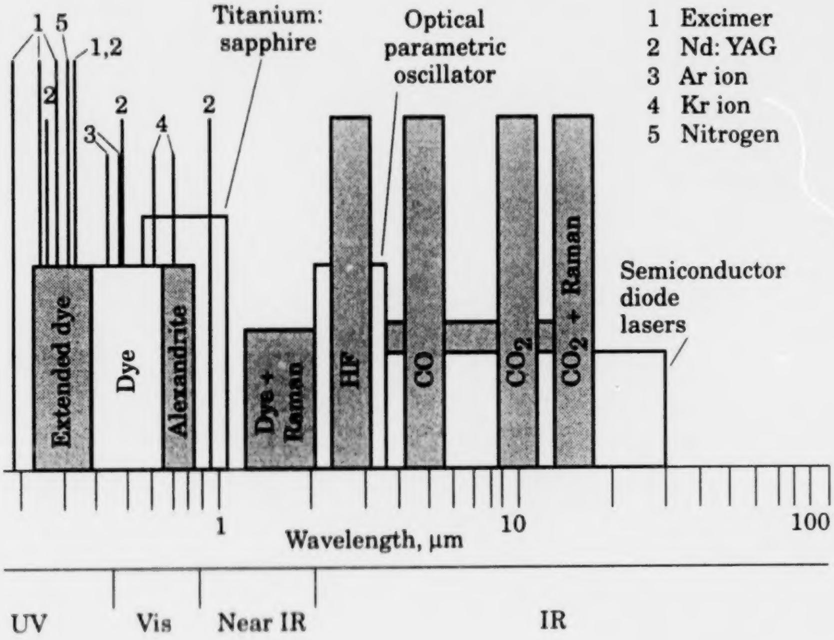
- **Transparency Masters**
  1. Wavelengths of Commercial Lasers (Andrews, 1986, p. 164; Woodin *et al.*, 1990)
  2. Laser Wavelength Coverage
  3. Solar Spectral Radial Distribution (Coyle *et al.*, 1982, p. 10)
  4. Fifteen Causes of Color (Nassau, 1983, p. 23)
  5. Causes of Color
  6. Excitation, Fluorescence, and Phosphorescence
  7. Word Search
- **Humor**

Wavelengths of Commercial Lasers

λ/nm	Laser	λ/nm	Laser	λ/nm	Laser
157	Fluorine	441.6	Helium-cadmium	534	Manganese
173.6	Ruby x 4	454.5	Argon	539.5	Xenon
193	Argon fluoride	457.7	Krypton	543.5	Helium-neon
222	Krypton chloride	457.9	Argon	568.2	Krypton
231.4	Ruby x 3	461.9	Krypton	578.2	Copper
248	Krypton fluoride	463.4	Krypton	595.6	Xenon
266	Neodymium x 4	465.8	Argon	628	Gold
308	Xenon chloride	468.0	Krypton	632.8	Helium-neon
325	Helium-cadmium	472.7	Argon	647.1	Krypton
333.6	Argon	476.2	Krypton	657.0	Krypton
337.1	Nitrogen	476.5	Argon	676.5	Krypton
347.2	Ruby x 2	476.6	Krypton	687.1	Krypton
350.7	Krypton	482.5	Krypton	694.3	Ruby
351	Xenon fluoride	484.7	Krypton	722	Lead
351.1	Argon	488.0	Argon	752.5	Krypton
353	Xenon fluoride	495.6	Xenon	799.3	Krypton
355	Neodymium x 3	496.5	Argon	904	Gallium arsenide
356.4	Krypton	501.7	Argon	1060	Nd:glass
363.8	Argon	510.5	Copper	1064	Nd:YAG
406.7	Krypton	514.5	Argon	1092.3	Argon
413.1	Krypton	520.8	Krypton	1152.3	Helium-neon
415.4	Krypton	528.7	Argon	1315	Iodine
428	Nitrogen	530.9	Krypton	1319	Nd:YAG
437.1	Argon	532	Neodymium x 2	3391	Helium-neon
				3508	Helium-xenon

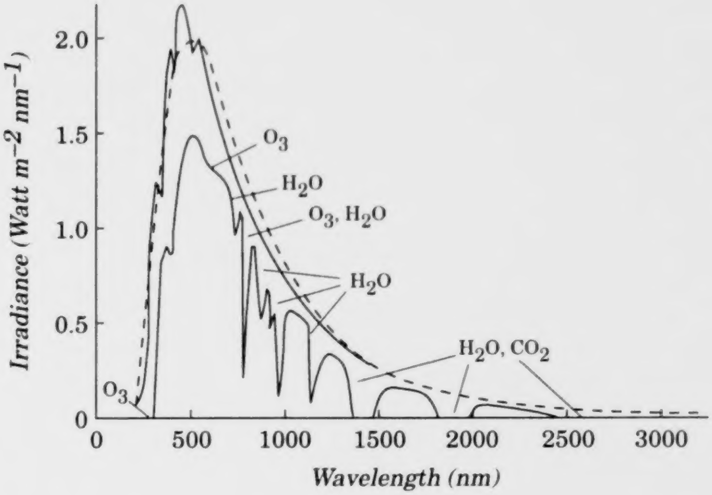
2608-3093 (21 lines)	Hydrogen fluoride	a 9160-9840 (45 lines)	Carbon dioxide
3493-4100 (37 lines)	Deuterium fluoride	10400-11040 (58 lines)	Nitrous oxide
5090-6130 (14 lines)	Carbon monoxide	a 10070-11020 (50 lines)	Carbon dioxide

Laser Wavelength Coverage



Solar Spectral Radial Distribution

Note that the detection limits of the human eye exactly match the maximum in this curve, indicating an evolutionary link between the development of vision and the environment.



Fifteen Causes of Color

Vibrations and Simple Excitations

1. **Incandescence** Flames, lamps, carbon arc, limelight
2. **Gas Excitations** Vapor lamps, lightning, auroras, some lasers
3. **Vibrations and Rotations** Water, ice, iodine, blue gas flame

Transitions Involving Ligand Field Effects

4. **Transition Metal Compounds** Turquoise, many pigments, some fluorescence, lasers, and phosphors
5. **Transition Metal Impurities** Ruby, emerald, red iron ore, some fluorescence and lasers

Transitions between Molecular Orbitals

6. **Organic Compounds** Most dyes, most biological colorations, some fluorescence and lasers
7. **Charge Transfer** Blue sapphire, magnetite, lapis lazuli, many pigments

Transitions Involving Energy Bands

8. **Metals** Copper, silver, gold, iron, brass, "ruby" glass
9. **Pure Semiconductors** Silicon, galena, cinnabar, diamond
10. **Doped or Activated Semiconductors** Blue and yellow diamond, light-emitting diodes, some lasers and phosphors
11. **Color Centers** Amethyst, smoky quartz, desert "amethyst" glass, some fluorescence and lasers.

Geometrical and Physical Optics

12. **Dispersive Refraction, Polarization, *etc.*** Rainbow, halos, sun dogs, green flash of sun, "fire" in gemstones
13. **Scattering** Blue sky, red sunset, blue moon, moonstone, Raman scattering, blue eyes and some other biological colors
14. **Interference** Oil slick on water, soap bubbles, coating on camera lenses, some biological colors
15. **Diffraction** Aureole, glory, diffraction gratings, opal, some biological colors, most liquid crystals.

Causes of Color*

Vibrations and Simple Excitations

- Incandescence
- Gas Excitations
- Vibrations & Rotations

Electronic Energy Levels in Ions or Molecules

- Transition Metal Compounds
- Transition Metal Impurities
- Organic Compounds
- Charge Transfer

Transitions Involving Energy Bands

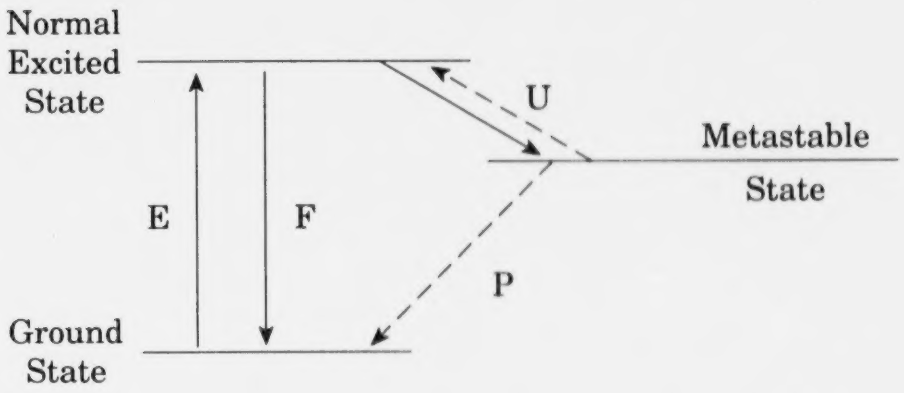
- Metals
- Pure Semiconductors
- Doped or Activated Semiconductors
- Color Centers

Geometical and Physical Optics

- Dispersive Refractions, Polarization, *etc.*
- Scattering
- Interference
- Diffraction

*Transparency 4 has more details.

Excitation, Fluorescence, and Phosphorescence


Normal
Excited
State
Ground
State
Metastable
State
E
F
U
P

E = Excitation

F = Fluorescence (allowed, fast)

P = Phosphorescence (forbidden, slow)

U = Unfavorable (slow)

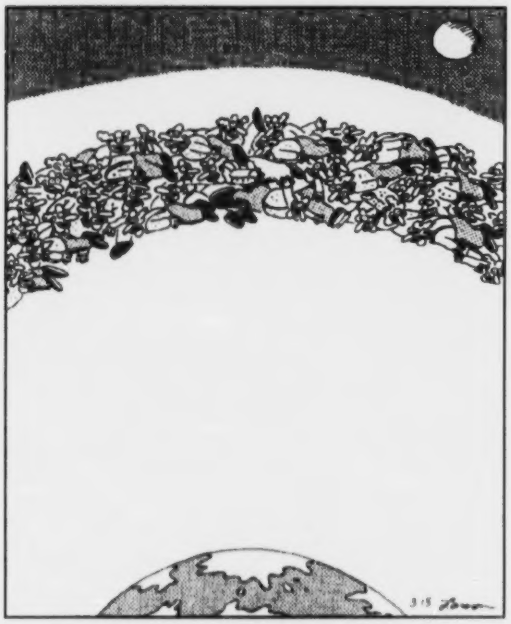
Word Search

T R I B O L U M I N E S C E N C E
C J U P P J U M M U L W Y C R O W
X E N Q H R O H P S O H P M H X A
O L E D K Q B M J K A U W Y J K V
E C N E C S E N I M U L I M E H C
R K G Q E B G D W Q G Y V L A P K
E C Q V T A E J V G J N Q E D D M
V I B U W R E S A L W R A F C J M
L K Q W A C H R O M O P H O R E S
I F Y R A N S Q I G O C H D V X C
S H F D R F T E L O I V A R T L U
P N J D O M L U G L R H Q D K E X
I D Y W I Z D E M M F K C N A L P

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. **Spectral region associated with molecular vibrations and heat.**
2. **Acronym for light amplification by stimulated emission of radiation.**
3. **Material that exhibits phosphorescence**
4. **Kind of yield that measures photochemical reaction efficiency.**
5. **Emission of light through chemical reaction.**
6. **Mechanically induced emission of light.**
7. **Another name for "black light."**
8. **Chemical groups that absorb light.**
9. **Halides of this metal are light sensitive.**
10. **ΔE = hv is called the Einstein-_______ relationship.**

THE FAR SIDE By GARY LARSON



The bozone layer: shielding the rest of the solar system from the Earth's harmful effects.

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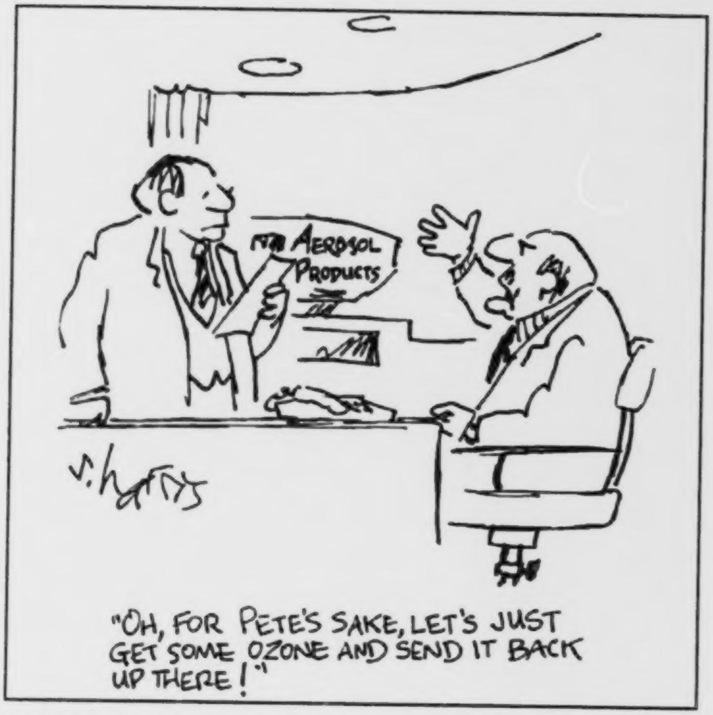
AEROSOL PRODUCTS
S. Harris
"OH, FOR PETE'S SAKE, LET'S JUST GET SOME OZONE AND SEND IT BACK UP THERE!"

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WHAT'S SO NEW ABOUT SOLAR ENERGY?
DAILY TABLET

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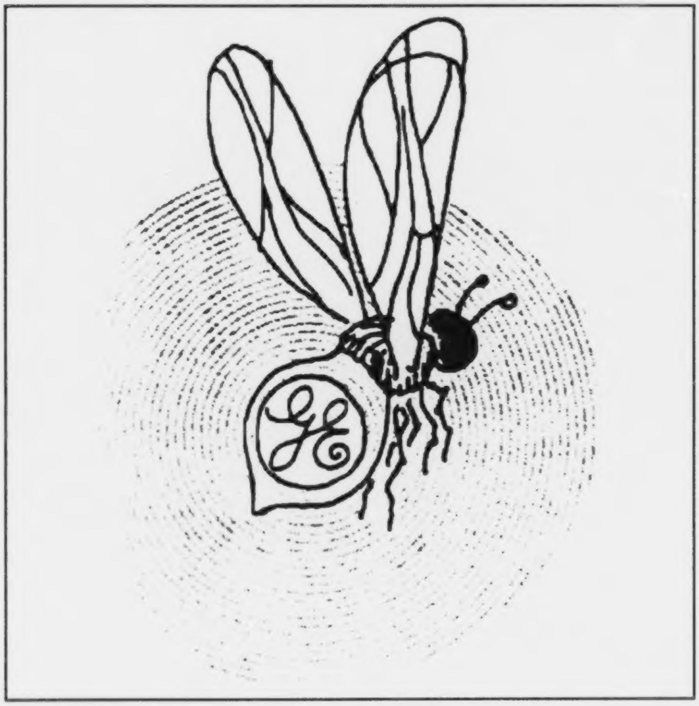

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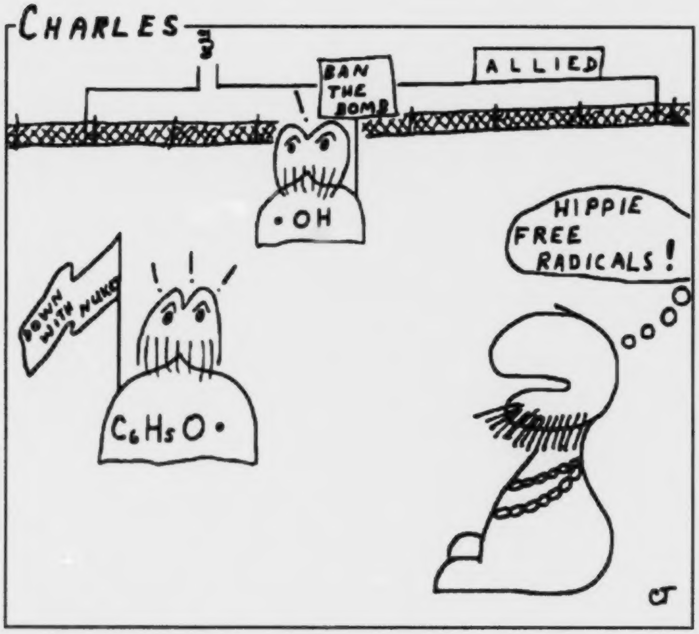
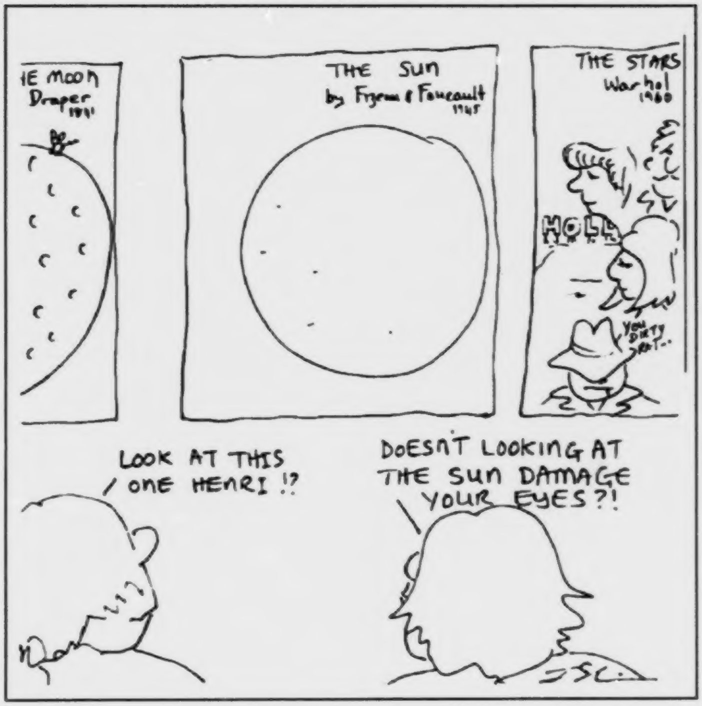
251


CHARLES
ALLIED
BAN THE BOMB
DOWN WITH NUKE
•OH
C_6H_5O•
HIPPIE FREE RADICALS!
CT

CHEM 13 NEWS, January 1984, p. 9. Reprinted with permission.


IE MOON
Draper
1841
THE SUN
by Fizeau & Foucault
1845
THE STARS
Warhol
1960
HOLL
YOU DIRTY RAT..
LOOK AT THIS ONE HENRI !?
DOESN'T LOOKING AT THE SUN DAMAGE YOUR EYES?!

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A SourceBook Module

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POLYMERS

ChemSource

Instructional Resources for Preservice and
Inservice Chemistry Teachers

Topic Overview

Content in a Nutshell

History is often described in terms of the materials employed in making useful items—the Stone Age, the Bronze Age, and the Iron Age. The 20th century is appropriately known as the Polymer Age, since most items we use daily are made from polymers.

Natural polymers have existed since the beginning of life itself—polymers such as DNA, RNA, proteins, cellulose, and starch. Over two thousand years ago, the Chinese developed paper making; Central American natives learned how to make rubber balls from the juice of certain trees. In the 1800s scientists and inventors began to convert natural polymers into products with useful properties. Charles Goodyear learned how to vulcanize rubber; Christian Schonbein changed cellulose into nitrocellulose; John Hyatt made celluloid from nitrocellulose and camphor; Louis Chardonnet produced artificial silk from nitrocellulose.

Large-scale production of purely synthetic polymers began in 1909 when Leo Baekeland developed the phenolic resin he called Bakelite. The polymer industry grew exponentially, and today it is almost impossible to list all the various polymers available. In the 1920s, the term "plastics" was used to describe these new materials. It is estimated that 75% of all chemists working today are in some field of polymers and plastics.

Simply described, polymers are giant molecules made up of long chains of repeating units. *Poly* means "many" and *mer* means "part." All polymers are made from small units, called **monomers**, joined in long chains by covalent bonds. For example, polyethylene is made from many ethylene molecules, and polystyrene is made from styrene molecules.

The process of polymerization can proceed in two general ways: addition and condensation, depending upon the types of monomers used. In **addition polymers** (chain-growth polymers), one monomer simply adds to the next, which adds to another, and another, *etc.* Chains of polyethylene can have 10,000-20,000 ethylene units. In **condensation polymers** (step-growth polymers), two different monomers are joined together with the elimination of a small molecule, like water. Nylon and polyester are examples of condensation polymers.

If the polymer chains are cross-linked, the product will usually be much more rigid and resistant to heat. **Thermoplastic polymers** can be heated and remolded many times, and **thermoset polymers** cannot be remolded once they have been originally shaped because they are cross-linked. Heat does not soften thermosets; they simply char if exposed to flame. Polyethylene and nylon are thermoplastics; Formica™ counter tops and epoxy putty are thermosets. Toothbrush handles are made of a thermoplastic polymer, while handles of kitchen utensils are fabricated from thermoset polymers.

Place in the Curriculum

This is an enrichment module. The material could be presented in conjunction with a unit on organic chemistry.

Central Concepts

1. **Polymers** are large molecules made up of long chains of repeating units called **monomers**.
2. The monomers are held together by covalent bonds.

3. Polymers are formed by two general types of reactions, addition (chain-growth polymers) and condensation (step-growth polymers).
4. Natural polymers include rubber, cellulose, starch, DNA, proteins, *etc.*
5. The physical properties of polymers are affected by cross-linking; therefore, thermoplastic and thermoset polymers have different physical properties.
6. We live in an age of polymers. Most products we use daily are polymers, and would be missed if eliminated. Disposal of some polymers is an environmental concern.

RELATED CONCEPTS

1. Covalent bonding
2. Basic organic chemistry (tetravalent carbon; single, double, and triple bonding; organic compounds containing nitrogen, halogens, and oxygen)
3. Catalysis
4. Hydrogen bonding

RELATED SKILLS

1. Writing structural formulas for organic compounds.
2. Good observational skills.
3. Ability to follow directions.

PERFORMANCE OBJECTIVES

After completing their study of polymers, students should be able to:

1. give the names of common polymers.
2. combine monomers to form addition and condensation polymers.
3. explain the difference between thermoplastic and thermoset polymers, and give examples of both.
4. explain the meaning of cross-linking.
5. discuss the importance of polymers in our lives.
6. discuss problems of solid waste disposal and recycling of polymers.

Concept/Skills Development

Laboratory Activity: Student Version

Activity 1: Thermoplastic and Thermoset Polymers

Introduction

Two general types of polymers and their properties are:

Type	Properties
Thermoplastic	Can be molded and remolded many times, using heat
Thermoset	Once set, cannot be remolded

Thermoplastic polymer chains are not connected to each other, while thermoset polymer chains have numerous cross-links that bond them into rigid networks.

Purpose

To observe properties of a thermoplastic polymer, Friendly Plastic™, and a thermoset polymer, epoxy putty.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Thermoplastic activity: You will be using boiling water; observe normal precautions. Do *not* make a bracelet or ring out of the Friendly Plastic™. Also, do *not* put the plastic in your mouth.
3. Thermoset activity: While working with the epoxy putty, do *not* put your fingers in your mouth or eyes. Clean your hands with rubbing alcohol when finished.

Procedure

Thermoplastic Activity

1. Pour about 300 mL tap water into a 400-mL beaker; bring the water to a boil. Then turn off the burner or hot plate.
2. Hold one end of a strip of Friendly Plastic™, and dip the other end into the hot water until it softens. Keep it away from the sides of the beaker. Remove it from the water and—with *wet* fingers—grasp the softened end (it will not burn you). Dip the other end in the water until it is also softened.
3. Remove the strip from the water bath, and quickly mold the Friendly Plastic™ into a design of your own choosing—*not a bracelet or ring!* If the polymer becomes too hard, dip it in hot water to soften it.
4. Let the material cool on the counter top.
5. Thoroughly wash your hands before leaving the laboratory.

Thermoset Activity

1. Obtain a 2-cm piece of epoxy putty from your teacher. Note that it is composed of two colored components.
2. Knead the putty until the two components are thoroughly mixed.

3. Quickly shape the putty into a design of your choosing.
4. Let the putty harden on a paper towel or piece of plastic bag. Do *not* let the putty harden on the counter top; it may be very difficult to remove.
5. As the putty is hardening, test its temperature with your finger tips. Does the sample change temperature? Describe what you observe.
6. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. In your words, describe the differences between thermoplastic and thermoset polymers.
2. List at least ten items of each type of polymer that you have used or seen.

Implications and Applications

1. Which of these two types of polymers could be more readily recycled?
2. In what applications would a thermoset polymer be most desirable?
3. Which polymer type would you choose for:
    a. Microwave dish
    b. A plastic bag
    c. Toothbrush handle
    d. Child's toy (doll or game)
    e. Dinnerware

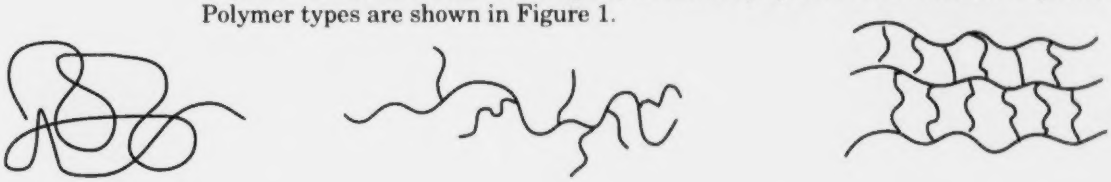
LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Thermoplastic and Thermoset Polymers

NOTE: Most of these activities require substances that are not normally in school stockrooms. Plan ahead!

Major Chemical Concept

The two general polymer types differ in properties because a thermoplastic polymer has only long, unconnected chains (randomly coiled), while a thermoset polymer has many cross-links, leading to a very rigid structure. Thermoplastic polymers can be heated and molded over and over again; thermoset polymers can be molded only once. Polymer types are shown in Figure 1.



Linear random coil

Branched polymer

Cross-linked polymer

Figure 1. Models of linear random coil, branched, and cross-linked polymers.

Level

Basic/general

Expected Student Background

Students should know:

1. how polymers are formed from monomers.
2. the terms thermoplastic and thermoset.
3. how thermoplastics and thermosets differ in structure.
4. everyday examples of thermoplastics and thermosets.

Time

40-50 min

Safety

Read the *Safety Considerations* in the *Student Version*. Emphasize that students must *not* make bracelets or rings from these polymers, since the plastic will harden and become very difficult to remove. Also, students must *not* put any of the polymers in their mouths because it will stick to metal such as orthodontic braces. Provide rubbing alcohol to clean hands after using the epoxy putty. *CAUTION: Rubbing alcohol is flammable.*

Materials (For 24 students working in pairs)

12 Sticks of Friendly Plastic™ (May be obtained in art supply stores or from AMACO Co., Indianapolis.)
12 Beakers, 400-mL
12 Burners or hot plates
1.5 Epoxy putty sticks (approximately 1.4 kg) or any similar product (found in the plumbing section of hardware stores)

Advance Preparation

Cut each epoxy putty stick into 1-in long pieces.

Pre-Laboratory Discussion

Review the structures of thermoplastic and thermoset polymers.

Demonstration

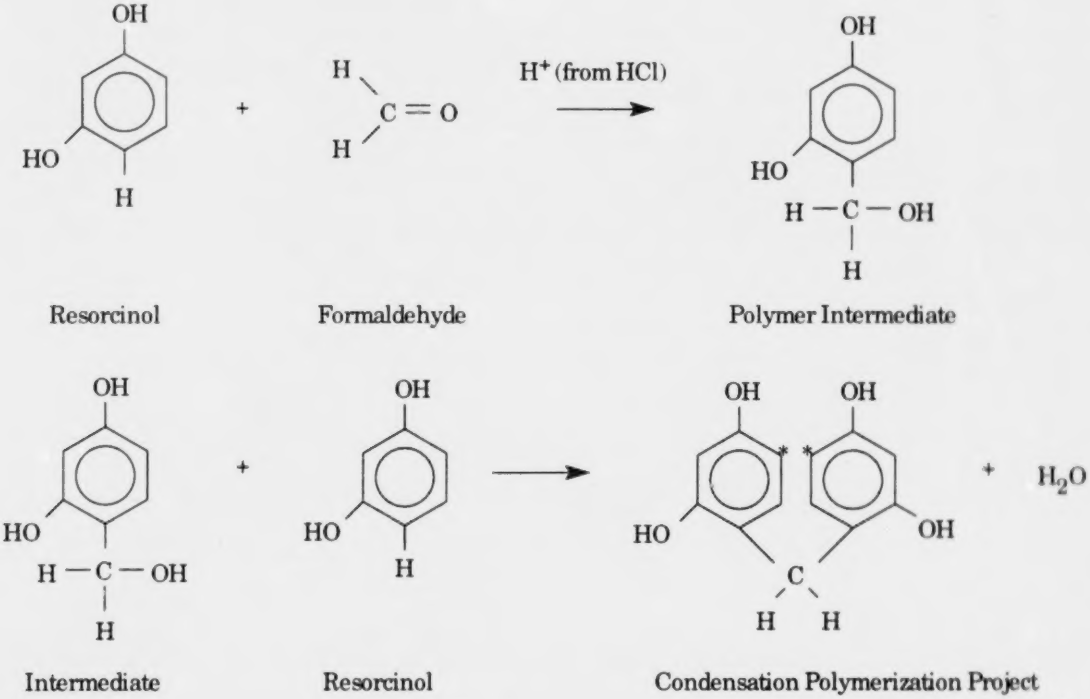
Boil 400 mL water in a 600-mL beaker. Place a clear plastic cup (which is not foam or recycle code #6 plastic) and the handle end of a kitchen tool such as a spatula into the boiling water. Let stand in the boiling water for 2-3 min. Remove the cup with tong and show that it can be molded into a new shape. Then remove the kitchen tool to show that the handle does *not* soften or bend. This represents two types of polymers.

How do the properties of these two kinds of polymers determine their uses? Elicit examples of items in students' homes that would be in each category.

Optional Demonstration

Making a thermoset polymer: *Take care in the use of resorcinol and formalin.* Both substances are toxic. Formalin is also an irritant.

1. Weigh approximately 2 g of resorcinol (1,3-dihydroxybenzene) into a 50-mL beaker. Add 3 mL 37% formaldehyde (formalin).
2. Place the beaker in a water bath; keep the water bath at 50 °C until all the resorcinol has dissolved. Stir while dissolving.
3. Remove the stirring rod; replace with a loop of copper wire.
4. Remove the beaker from the water bath. Add dropwise 6-8 drops concentrated hydrochloric acid, HCl, until a reaction occurs. *Caution: Reaction is very exothermic!*
5. Pull the cooled product from the beaker with the wire. Examine it. Wash hands if you handled the product.



OH
HO
H
+
H
C = O
H
H⁺ (from HCl)
OH
HO
H — C — OH
H
Resorcinol
Formaldehyde
Polymer Intermediate
OH
HO
H — C — OH
H
+
OH
HO
H
OH
OH
*
*
HO
OH
C
H
H
+
H₂O
Intermediate
Resorcinol
Condensation Polymerization Project

Condensation in this way will continue at the ortho-benzene positions (adjacent to the –OH) to produce long chain-like molecules; reaction at the para-benzene positions (directly opposite the –OH) results in cross-linking of the long chains to yield a rigid thermoset polymer. [These positions are marked with asterisks in the structure.] In commercial processes, the extent of cross-linking can be controlled to suit the requirements of the final product.

Teacher-Student Interaction

Thermoplastic Activity

1. Why is it necessary to heat Friendly Plastic™ in order to shape it? *[To give the polymer chains enough energy to slip and slide over one another.]*
2. Is this an exothermic or endothermic change? *[Endothermic]*

   Is this a physical or chemical change? *[Physical]*
3. How many times can this material be remolded? *[Many times, as long as the polymer chains stay intact.]*
4. How could this material be used? *[For art sculptures and in industrial molds as for temporary support for small objects.]*

Thermoset Activity

1. Why are the two epoxy putty components separated? *[To prevent premature reaction.]*
2. Why are two components kneaded? *[Because the reaction will not take place without mixing the components.]*
3. As the epoxy putty sets, does it feel hot or cold? *[Hot]*
4. Is this an exothermic or endothermic reaction? *[Exothermic]* Is it a physical change or chemical reaction? *[Chemical reaction]*
5. Could you remold this material if you do not like the first effort? Explain. *[No. This chemical reaction yields a rigid, new product by cross-linking.]*

Anticipated Student Results

Thermoplastic Activity

Students will be able to mold a piece of low-melting Friendly Plastic™, and discover that it can be molded repeatedly. They should be able to list a number of household items made from thermoplastic polymers, and conclude that thermoplastics are more easily recycled.

Thermoset Activity

Students will be able to mold a piece of epoxy putty and observe that it gives off heat as it sets. They should be able to draw a diagram of the cross-linking process. They should give examples of household items made from thermoset polymers. They should be able to predict what will happen when a thermoset polymer is placed in boiling water.

Answers to Data Analysis and Concept Development

1. Thermoplastics can be repeatedly molded, but thermosets cannot (see Figure 1 for molecular-scale models).
2. Items made of polyethylene, nylon, polyester, polyvinyl chloride, polystyrene are thermoplastics. Examples include clothing, and plastic packaging material, such as bottles and bags. Thermosets are used in bowling balls, Formica™ table tops, auto body parts, and kitchen utensils.

Answers to Implications and Applications

1. Thermoplastic polymers can be recycled; all that is necessary is to melt them again. (You should point out that polymers have to be *sorted*, however, and that the plastics industry has initiated a voluntary labeling program for all thermoplastics. See *Appendix* for plastic bottles codes.)
2. Where strength and heat resistance are important.
3. (a) Thermoset; (b) Thermoplastic; (c) Thermoplastic; (d) Thermoplastic; (e) Thermoset

Post-Laboratory Activities

1. Students may have made jewelry items (pendants or earrings) from the Friendly Plastic™. Holes can easily be drilled in this material. Try drilling a hole in a piece of Friendly Plastic™; compare with drilling a hole in Epoxy Putty. Is there any difference?
2. Both polymer types can be painted with acrylic paints or permanent colored markers.

Extensions

1. Test the flammability of a small piece of Friendly Plastic™. *[It is flammable, but it melts first.]*
2. Test the resistance of epoxy putty to a burner flame. *[It is nonflammable. It doesn't melt; it chars.]*
3. How does environmental temperature affect the epoxy putty's setting time? This is a good homework activity to test student observations. Have students observe the sample in a freezer, a refrigerator, at room temperature, and in a warm incubator. Also, try to set it under water.

Assessing Laboratory Learning

1. Draw "pictures in the mind" of how polymer chains in the Friendly Plastic™ look before and after being molded. Imagine spaghetti (see Figure 1).
2. Draw "pictures in the mind" of how the polymer chains in the epoxy putty might look, both before and after molding. Imagine "floppy" ladders (see Figure 1).

References

Thermoset polymer demonstration: Ferguson, Schmuckler, Caro, and Siegelman. (1973). *Laboratory Investigations in Chemistry*, Teacher's Edition. Silver Burdett. (Also contains information on variations in the activity.)

Epoxy molecular structure: *Encyclopedia of Science and Technology, Vol. 10*. (1974). New York, NY: McGraw-Hill.

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: Preparation and Properties of a Cross-Linking Polymer

Introduction

Solutions of the polymer polyvinyl alcohol (PVA) and polyvinyl acetate can be made into gels by addition of a borax solution. Borax cross-links the polymer chains. In this activity you will investigate interesting properties of gels. "Slime" is a polyvinyl alcohol polymer. "Gluep" is a polyvinyl acetate polymer. Your teacher will indicate which you will prepare.

Purpose

To observe the effects of cross-linking polyvinyl alcohol and polyvinyl acetate chains with borax. (*OPTIONAL: To observe the reversal of the cross-linking process.*)

Safety

1. Wear protective goggles throughout the laboratory activity.
2. The polyvinyl alcohol and borax solutions are nontoxic, but wash your hands when finished. Keep the gel in a plastic bag; keep off clothes and carpets.

Procedure

Part I. Making "Slime"

1. Obtain 20 mL PVA solution in a disposable cup.
2. Obtain a wooden stirring stick—a glass stirring rod does not work as well—and stir the PVA solution. Note its viscosity. Is it more or less viscous than water?
3. If desired, add one drop food color to the PVA. Stir well.
4. Measure 5 mL borax solution in a 10-mL graduated cylinder, or use a plastic teaspoon, which holds approximately 5 mL.
5. Pour the borax solution into the PVA. Stir vigorously until the gelling is complete.
6. Scrape the gel into your hands. Pat and knead it thoroughly to completely mix the components.
7. Test its properties as directed by your teacher (density, viscosity, malleability, resistance to pressure).
8. Take the sample home in a zip-closure plastic bag.
9. Thoroughly wash your hands before leaving the laboratory and each time you handle "slime."

Alternate Part I. Making "Gluep"

1. Mix white glue 50:50 with water. Stir well. (Your teacher may have done this.)
2. Measure two tablespoons (30 mL) glue mixture into a cup.
3. Add color and stir well with a stick.
4. Add two teaspoons (10 mL) borax solution and stir vigorously. (The "gluep" will cling around the stick and can be pulled out, usually in one blob, without much being left in the cup.)

5. Knead the "gluep" well, and observe its properties as before and compare to properties of slime.
6. Store "gluep" in a tightly sealed zip-closure bag. Throw away when it gets moldy.
7. Thoroughly wash your hands before leaving the laboratory and each time you handle "gluep."

Part I. Implications and Applications

1. Is covalent or hydrogen bonding responsible for the cross-linking?
2. Why is borax [$B(OH)_4^-$ is the active ingredient], an effective cross-linker?
3. Draw a representation of the polymer chains:
   a. before cross-linking.
   b. after cross-linking.

   (Be sure to include some water molecules in your drawing; after all, both solutions are 96% water!)

Part II (Optional). Effect of Acid and Base on "Slime"

8. Place a dime-sized piece of slime on a watch glass or Petri dish. Add 2 M HCl dropwise, stirring well with the stick after each drop. When a change is noticed, record your observations.
9. Add 2 M NaOH dropwise to the same sample, stirring well with the stick after each drop. When a change is noticed, record your observations. (Repeat the procedure, if you wish.)
10. Thoroughly wash your hands before leaving the laboratory.

Part II. Data Analysis and Concept Development

1. How many drops of HCl were necessary to "break the gel?"
2. How many drops of NaOH were required to re-gel the polymer?
3. Write the equation for the reaction of HCl with NaOH.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: Preparation and Properties of a Cross-Linking Polymer

Major Chemical Concept

Borax (sodium borate, $Na_2B_4O_7 \cdot 10H_2O$) can be used to cross-link chains of polyvinyl alcohol polymer in water. The resulting network of cross-linked chains produces a gel that traps water molecules. This gel has fascinating properties when handled, squeezed, stirred, poured, or stretched. It can be used to demonstrate acid-base reactions and equilibrium.

Level

Appropriate for students at basic, general and honors levels.

Expected Student Background

Students should know:

1. how polymers are formed from monomers.
2. the functional organic alcohol group, –OH.
3. the concept of hydrogen bonding.

Optional:

4. acid-base reactions.
5. the concept of equilibrium,

Time

40-50 min

Safety

Read the *Safety Considerations* in the *Student Version*. The polyvinyl alcohol (PVA) solution, borax solution and gel are not toxic. However, students should be cautioned not to make a mess with the gel and not to let young children play with it. Mention that if the gel develops mold, it should be discarded.

Materials **(For 24 students working in pairs)**

Making "Slime"

- 4% Polyvinyl alcohol, PVA, 240 mL (see *Advance Preparation*)
- 4% Borax, 100 mL (4.0 g household Borax diluted to 100 mL)
- 12 Plastic or paper cups
- 12 Plastic teaspoons or 10-mL graduated cylinders
- 12 Popsicle sticks or tongue depressors
- Food coloring
- 24 Zip-closure plastic bags
- 12 Graduated cylinders, 50-mL
- 12 Watch glasses or Petri dishes

Making "Gluep"

- Elmer's™ (or any other white) glue, 200 mL
- 4% Borax, 100 mL (4.0 g household Borax diluted to 100 mL)
- Food coloring or tempera colors
- 12 Stir sticks (craft sticks)
- 12 Plastic or paper cups, 6 or 10 oz.
- 24 Zip-closure plastic bags
- 12 Teaspoon measures
- 12 Tablespoon measures

Effect of Acid and Base on PVA (Optional)

2 M Hydrochloric acid, HCl, 100 mL (16 mL conc. HCl diluted to 100 mL)
2 M Sodium hydroxide, NaOH, 100 mL (8 g NaOH per 100 mL solution)

Advance Preparation

For "Slime"

PVA Solution: Use 99% hydrolyzed polyvinyl alcohol powder. (Can be ordered from Fisher 91-92 EK 153 9709.) To prepare a 4% solution, weigh 40 g 99% hydrolyzed polyvinyl alcohol. Gradually heat one liter of tap water. When it reaches 40-50 °C, gently sprinkle the powder into the water while constantly stirring (or use a magnetic stirrer). Keep heating gently and stirring until about 95 °C is reached; do not overheat or allow to boil. The solution will look cloudy until it reaches 85-90 °C; then it will clear. Remove from heat; cover with aluminum foil, and allow to cool overnight. Store in plastic bottles. It will keep for a long time.

For "Gluep"

50:50 Glue Solution: To make 50:50 glue solution, empty the bottle of glue into a container, letting it drain well. Then fill the glue bottle with water, shake, and empty into the container. (It is easier to make this ahead of time than to have students prepare the solution.)

Prepare other solutions in advance as well. A mixture of liquid starch and glue also works, but you need to determine the appropriate amounts of each by trial and error.

Pre-Laboratory Discussion

Show students the dry PVA powder and the 4% Borax solution. Demonstrate the solution's viscosity by pouring from one beaker to another. Explain how PVA and borax solutions were made. Show the structural formula of polymer chains and borate ion. In water, the borax hydrolyzes to form a borate-boric acid buffer system:

$$B(OH)_3 + 2\,H_2O \rightleftharpoons B(OH)_4^- + H_3O^+$$

The $B(OH)_4^-$ ion is believed to cross-link the polymer chains as shown in Figure 2.

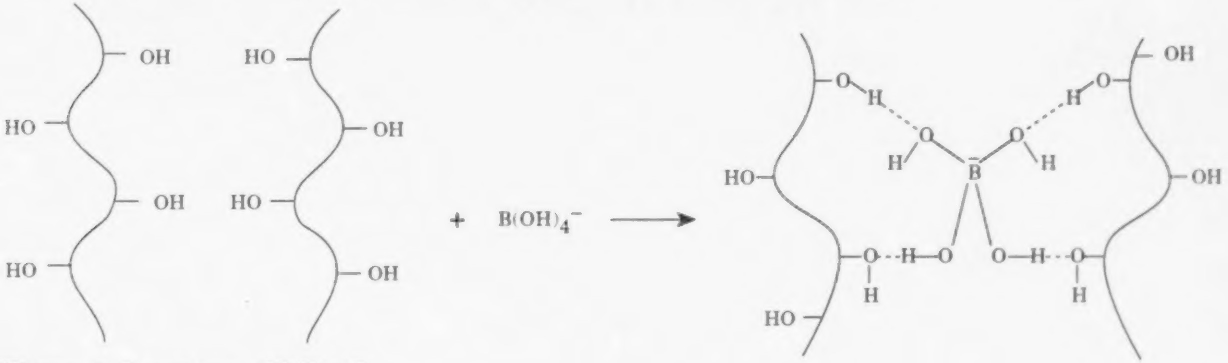


Figure 2. Structure of "slime"

Demonstration: "People-Polymers"

1. Arrange four students in a line holding hands; put four other students in a second line behind the first line. These represent the PVA chains. Can they move freely?

2. Ask two other students to walk between the lines, and take hold of a student on each side. These students represent the borate ions. Can the PVA chains move freely now? If there were water molecules between the chains, what would happen to them? (They would be trapped in a three-dimensional network, which we call a gel.)

Teacher-Student Interaction

Part I

1. Why is PVA solution so viscous? *[Because long polymer chains cannot move freely.]*
2. Try to imagine the many long chains of molecules, being pushed around as you stir it. It is sort of like a bucket full of eels.
3. After the cross-linking, why does the gel have to be kneaded thoroughly? *[To provide complete mixing.]*
4. How is the gel like a liquid? *[It flows, has no definite structure, takes any shape.]* How is it like a solid? *[It can be broken with a sharp pull; it bounces.]*
5. Be sure students understand what was seen: PVA chains cross-link with the $B(OH)_4^-$ ion, forming a gel that traps water molecules.

Part II

6. Why does acid destroy the properties of the gel? *[Hydronium ions break hydrogen bonds between chains.]*
7. How have the properties changed after addition of NaOH? *[It again becomes a gel—stretches, bounces, drips.]*
8. Why does base restore the gel's properties? *[It neutralizes hydronium ions, permitting cross-linking to reoccur.]*

Anticipated Student Results

Students produce a gel (colorful—if food coloring has been added) that has interesting properties: it can be dripped, stretched, patted, rolled into a ball, bounced on a *clean* surface, *etc.* Students should be expected to observe these properties and record them. Students will notice how rapidly the cross-linking takes place. Encourage them to ask the question, "If the borax solution is diluted to 1% or 2%, will the gelling occur as fast?" Then try it!

Optional Activity

When one drop of 2 M HCl solution is placed on a dime-sized piece of gel, the gel starts to dissolve in a few seconds. Adding one or two drops of 2 M NaOH solution restores the original properties of the gel.

Answers to Part I. Implications and Applications

1. Cross-linking occurs by hydrogen bonding, not covalent bonding.
2. Because –OH groups of the borate ion can hydrogen bond with –OH groups of PVA polymer chains (see Figure 3). Hydrogen bonding is so weak that this could account for the ease of stretching and dripping, and the other observed properties.

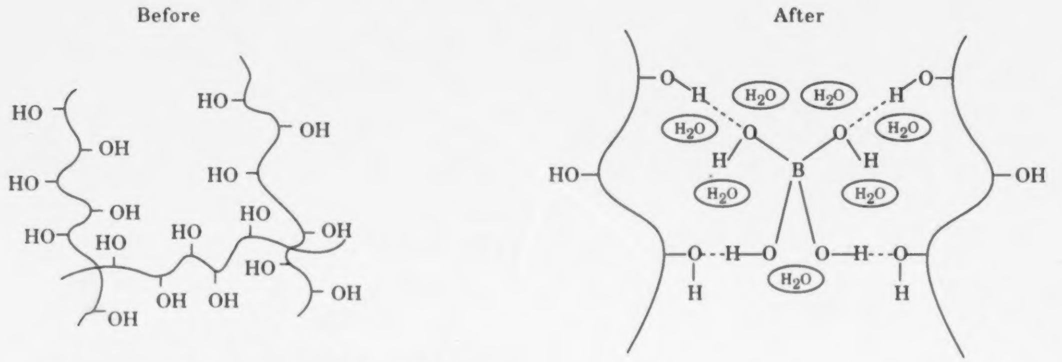


Figure 3. Hydrogen bonding in slime.

Answers to Part II. Data Analysis and Concept Development

1. (Observation)
2. One more drop than drops of HCl
3. $NaOH + HCl \rightarrow NaCl + HOH$

Post-Laboratory Discussion

1. Ask students to list the properties of PVA solution before and after gelling. *[Before gelling: syrupy, liquid. After gelling: if pulled suddenly will break, malleable, semi-solid.]*
2. How does hydrogen bonding explain the gel's properties (that is, how it can be squeezed, dripped, stretched, *etc.*)? *[Hydrogen bonds can be more easily broken and reformed; covalent bonds are too strong to cause these properties.]*
3. How do you explain the effect of HCl and NaOH, in terms of equilibrium?

$$\textit{[gel} \underset{\text{NaOH}}{\overset{\text{HCl}}{\rightleftharpoons}} \textit{liquid]}$$

4. Can you think of other uses for the gel other than a toy? *[Special effects in motion pictures,* e.g. *people's faces melting in "Raiders of the Lost Ark."]*

Extensions

1. Allow some gel to dry in air, and observe its properties *[becomes brittle and curls around the edge]*. Can it be rehydrated? *[No]*
2. Let students color some clear gel with fluorescent highlighter pens. Then observe under ultraviolet light. (*CAUTION*: Protect eyes from ultraviolet light!). *[Glows brilliantly!]*
3. Try using 2% PVA solution and 2% borax solution (or other proportions). *[No satisfactory gel is formed.]*
4. Will other acids and bases have the same effect as HCl and NaOH? *[Yes]*

Assessing Laboratory Learning

1. Draw "pictures in the mind" of the polymer before and after gelling. Include some water molecules in your picture.

2. Describe the viscosity of the PVA solution before and after gelling. *[Higher after gelling.]* Can it still be poured? *[Yes. It will run through the funnel.]*
3. In this activity, why is it not critical to use "exactly" 20 mL of PVA solution and "exactly" 4 mL of the borax solution? Why is it important in other activities to be *very* exact in your measurements? *[The extent of cross-linking depends on the amount of borax. The properties of "slime" depend on varying the number of cross-links.]*

References

Casassa, E. Z., Sarquis, A. M., and Van Dyke, C. H. (1986). The gelation of polyvinyl alcohol with borax. *Journal of Chemical Education, 63,* 57.

Sarquis, A. M. (1986). Dramatization of polymeric bonding using slime. *Journal of Chemical Education, 63,* 60.

CHEM FAX. Polyvinyl alcohol, preparation of a 4% solution for the 'slime' activity. Flinn Scientific. (131 Flinn Street, Batavia, IL 60510. Call (708) 879-6900 to be put on their mailing list.)

DEMONSTRATIONS

Caution: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Making Nylon Rope

Purpose

To prepare a condensation polymer and demonstrate how it grows.

Safety

Do not let any reagent touch skin. Sodium hydroxide is caustic and causes burns. This demonstration should be performed in well-ventilated areas.

Materials

NOTE: All of these solutions may be purchased from Flinn Scientific or other suppliers. Because of difficulty in preparation these commercial solutions should be used.

5% Hexamethylenediamine (1,6-hexanediamine); 5 mL hexamethylenediamine in 95 mL H_2O

20% Sodium hydroxide solution (NaOH); 20 g NaOH in 80 mL H_2O

4% Sebacoyl chloride in hexane (or 5% adipoyl chloride in cyclohexane); 4 g sebacoyl chloride in 96 mL hexane

Beaker, 50-mL

Procedure

Pour 10 mL 5% aqueous hexamethylenediamine (1,6-hexanediamine) into a 50-mL beaker. Add 10 drops 20% NaOH. Tilt the beaker and carefully pour 10 mL 4% sebacoyl chloride down the side of the beaker. Two layers will form with a polymer film between them.

Prepare a 15-cm piece of copper wire with a bend at one end. Hook the polymer film in the center, and raise slowly so nylon forms continuously as a rope. This can be many feet long. Break the rope by pulling faster. Rinse the rope several times with water and dry on a paper towel.

Demonstration 2: Plastic Packaging "Peanuts" in Acetone

Purpose

To illustrate properties of a plastic product.

Materials

Packaging "peanuts"
Acetone

Safety

CAUTION: Acetone is flammable. Acetone use is discouraged for safety reasons; however, some fingernail polish removers contain acetone and may be used, but do not work as well because they are too dilute. Test any solutions before demonstrating.

Procedure

A small volume of acetone softens a very large number of these packaging peanuts because they are air-filled foam. Using a large beaker, add the packaging peanuts to acetone while swirling. You may also soften a Styrofoam™ cup in acetone.

Demonstration 3: Super Slurper

Purpose

To illustrate properties of a plastic product.

Materials

Sodium polyacrylate (Super Slurper™, available from Flinn Scientific Co.)
Table salt (Sodium chloride, NaCl)
Disposable diaper

Procedure

1. Sprinkle about 1/2 teaspoon of the Super Slurper into 1/2 cup water. Observe the reaction. Try to pour it out of the cup. Take out a piece and try to squeeze out the water.
2. Put another piece of the gel into a separate cup and shake some table salt (NaCl) on top. Wait a few minutes and notice the consistency. Salt has a greater affinity for water than does the polymer and can break down the gel.
3. Using the material in a disposable diaper show the properties of this compound. The powder can be removed by cutting the diaper into 10 cm^2 pieces. Place the pieces in a plastic bag, close the bag and shake out the powder. This powder may then be treated as Super Slurper.

Remarks

Super slurpers, originally produced from starch and acrylonitrile, were invented at the Department of Agriculture Research Laboratory at Peoria, IL; however, the polyacrylate polymer is more efficient.

Demonstration 4: Properties of Natural Polymers

Purpose

To illustrate properties of natural polymers.

Materials

Latex rubber samples (Flinn Scientific)
Vinegar
Water
Petri dish

Safety

Sample is dissolved in ammonia. Smell carefully and do not spill.

Procedure

Pour a small amount of latex into the Petri dish and allow to dry. Notice its elasticity and other properties. Add some latex to water. Have a student cautiously smell the latex and identify the base in which it is dissolved (NH_3). Addition of vinegar will coagulate the rubber. Make a ball with the neutralized latex by pouring a small amount of latex into vinegar and hold product under water while kneading to form a ball.

Demonstration 5: Properties of Polymers

Purpose

To illustrate the properties of polymers.

Materials

Zip-closure plastic bag
Pencils

Procedure

Fill a plastic bag with water and seal. Push very sharp pencils through the bag. You are pushing apart the chains of the polymer as you put in the pencils. Pull out the pencils over a sink, and the water will gush out. Practice first.

Demonstration 6: Making Polyurethane Foam

Purpose

To make a synthetic polymer.

Safety

Isocyanates are toxic. This demonstration must be performed in a well-ventilated area. It is important that hardened foam not be handled for 24 hours to evaporate excess isocyanate. Store in a fume hood.

Materials (Available at a hobby store, from Flinn Scientific, or Edmund Scientific)

Polyether polyol, 50 mL
Poly-functional isocyanate, 50 mL
Container (paper cup, beaker or flat bottomed ice cream cone)
Wooden stir stick

Procedure

Mix equal volumes of the two components; stir well with wooden stirrer. As the reaction occurs, carbon dioxide gas is released and makes the polymer rise. The two reactants may be mixed in a beaker, a paper cup, or in a flat bottomed ice cream cone. This exothermic reaction illustrates the formation of a rigid, cross-linked material.

Demonstration 7: Polymers Have Different Directions

Purpose

To study the properties of polymers.

Materials

Tinted plastic bags
Polarized sheets
Overhead projector
Saran™ wrap

Procedure

Using tinted plastic bags show that polymers have different directions. Look through bags in a strong light. Cut rectangular strips lengthwise and crosswise. Stretch slowly and observe the different ways they separate. Stretch kitchen wrap (other than Saran™ wrap) between polarized sheets on an overhead projector. Notice the colors produced. Test the cellophane from cigarette packages in the same manner.

Demonstration 8: "Moon Blob"

Purpose

To illustrate the properties of a synthetic polymer.

Materials

"Moon Blob" (Edmund Scientific, Barrington, NJ; (609) 573-6250
Beakers

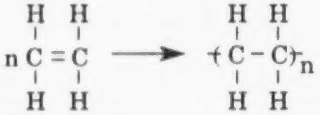
Procedure

Mix dry powder with specified amount of water. Mix thoroughly. Pour the "Moon Blob" (polyethylene oxide) slowly from one beaker to another and note its properties. It will not break apart.

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. Name some natural polymers. *[Cellulose, proteins, rubber, starch]*
2. Name some synthetic polymers. *[Nylon, orlon, rayon, dacron, etc.]*
3. What is a condensation (step-growth) polymer? Write an equation for the formation of a condensation polymer. *[A polymer formed from two different monomers that are joined together with the elimination of a small molecule. Nylon is a condensation polymer.]*
4. What is an addition (chain-growth) polymer? Write an equation for the formation of an addition polymer. *[A polymer formed from the addition of one monomer to another monomer without the loss of atoms. Polyethylene is an addition polymer formed as follows:]*



$$n\,\underset{H}{\overset{H}{C}}=\underset{H}{\overset{H}{C}} \longrightarrow \left(\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}\right)_n$$

5. Describe crosslinking. What is the difference between the bonds formed in crosslinking and those formed in polymer synthesis? *[A process that makes a polymer more rigid and resistant to heat. As shown in Figure 3, the cross-links in slime are hydrogen bonds. In the case of butadiene, the vulcanization process with sulfur cross-links the polybutadiene chains with covalent bonds as shown in Figure 4. (See also* Group Discussion—*additional suggestions about vulcanization.)]*

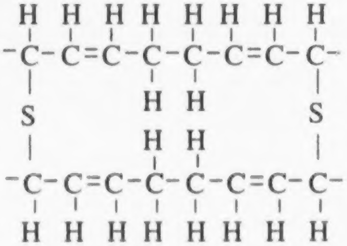


Figure 4. Vulcanized rubber.

6. What is the difference between thermoset and thermoplastic polymers? *[A thermoset polymer cannot be remolded; a thermoplastic polymer can be heated and remolded.]*

Counterintuitive Examples and Common Student Misconceptions

1. **"All polymers are products of chemical industries."**

   Rubber, starch, cellulose, proteins, and fibers like cotton and linen are all examples of natural polymers. These products are made in nature by chemical reactions.

2. **"All polymers are the same substance."**

   Compare the reaction of a polystyrene and a polyethylene cup to heat. Even all plastic kitchen wraps are not the same.

3. **"Automobiles are mostly metal."**

   In order to reduce mass and increase strength, modern automobiles contain a large proportion of polymeric materials in bumpers, steering wheels, upholstery, tires, windshields, paint, adhesives, *etc.* Some auto bodies, such as the Corvette™, are constructed of fiberglass.

4. **" Paper degrades under all conditions."**

   In landfills, where paper is covered by other waste, paper degrades very slowly. Urban archaeologists can date portions of landfills by noting the dates on newspapers and telephone books buried there.

5. **"Some plastic garbage bags are degradable."**

   "Biodegradable" bags contain cornstarch. Microorganisms can break down the bags into small pieces that degrade more quickly. However, landfills do not have the right conditions for this to occur, and the time required for degradation is many years.

Metaphors and Analogies

1. Polymers may be pictured as paper clip "chains," pop beads, or coupled freight train cars. Draw polymers starting with monomers, then chains, then cross-linked chains.
2. Play the polymer game. Have five to ten students form a line, standing about one arm's length apart with hands clasped in front of their chests. Each student represents an ethylene molecule. Act as the initiator, using a chalkboard eraser to represent a free radical. Unclasp the hands of a student on one end of the line (break the double bond), and hand the student the eraser. The student's free hand represents an unpaired electron, and the student now represents a free radical. This student unclasps the hands of the next student holding the closer hand and causing the second student to have an "unpaired electron." The chain is terminated by placing another eraser in the hand of the student on the other end [See Tarazona, M. P., and Satz, E. (1990). Happy polymer party. *Journal of Chemical Education, 64,* 238.]
3. Magnetic marbles on the overhead are a good illustration of chain formation. Pop beads may also be used.

Pictures in the Mind

Polymers may be pictured using analogies such as paper clip "chains," pop beads, or coupled freight train cars.

Group Discussion

This module lends itself well to discussion activities. Students could bring samples of plastics and discuss their uses (*e.g.*, empty margarine bowls, 2-L pop bottles, 1-gal milk containers, caulking compounds, insulation, super glue). They might discuss their daily life without plastic products. (See also introductory activity in *Appendix*.)

The problem of solid waste disposal and recycling could be discussed. Current periodicals are excellent sources of information. Students can prepare posters, give group reports or write papers to coordinate with Writing Across the Curriculum program (see *References* for sources of information or consult the school librarian).

Additional suggestions for discussion and research:

1. What are the advantages and disadvantages of using plastic containers rather than paper, metal or glass containers?
2. Consult local newspapers and discuss your community's efforts to recycle plastics and other materials.
3. General Electric Company has made a polyethylene plastic ice rink. After consulting magazine articles, comment on the use of this material and how it has changed hockey.
4. How did Astroturf™ change football?
5. Canoes made of Kevlar™ are very lightweight and strong. Discuss these with a local dealer and give a class report.
6. What are the advantages and disadvantages of using polyester or rayon blends as clothing?
7. Discuss the uses of plastics in medicine. Examples include pacemaker and hearing aid cases, balloons in angioplasty, replacement joints, and plastic surgery implants.
8. Vulcanization is the treatment of rubber with sulfur in the presence of a catalyst. Compare the properties of natural rubber, vulcanized rubber and synthetic rubber. Are tires made of natural rubber now? If not, why not? Natural rubber is *cis*-poly(1,3-butadiene), isoprene. The *trans* form produces gutta-percha, a harder, more brittle compound. Both are extracted from trees. Some synthetic rubber is also polyisoprene, but neoprene is 2-chloro-1,3-butadiene polymer. In vulcanization, sulfur forms cross-links between polymer chains to make the natural or synthetic rubber harder.
9. Discuss water absorbing polymers such as polyacrylamide, which is a soil additive, and polyacrylate, which is used in diapers.
10. Nylon was first used as a material for toothbrushes, and soon revolutionized the hosiery industry. During World War II, the U. S. government reserved all nylon for the war effort to make parachutes. Nylon hosiery was introduced at the New York World's Fair in 1939, but was not manufactured again until 1946.
11. Recycled plastics are used to make "bricks" for use in gardens and parks, and for benches in public areas. Identify parks or playgrounds in your community containing these plastic products. These products are more likely to be used outside because of the danger of harmful emissions being released during a fire.

TIPS FOR THE TEACHER

Language of Chemistry

1. **Glossary**

   addition polymer (chain growth polymer) polymer made by the addition of one monomer to another without the loss of atoms.

   biodegradable capable of being degraded by microorganisms.

   condensation polymer (step-growth polymer) polymer in which two different monomers are joined together with the elimination of a small molecule, like water.

   cross linking two polymer chains linked together by bonding.

   gel semi-rigid colloidal dispersion of a solid in a liquid.

   monomer small unit making up polymers.

   polyfunctional containing more than one functional group.

   polymer large molecule made up of repeating units called monomers.

   thermoplastic polymer polymer that can be heated and remolded many times.

   thermoset polymer polymer that cannot be remolded after being shaped.

2. **Polymers** are "giant molecules" formed from repeating units (*poly* = many and *mer* = parts). The individual units are called **monomers**; thousands of these are linked together to form a polymer molecule. "Giant molecules" have high molecular weight. For example, natural rubber has a molecular weight of 12,000; starch, 40,000; and proteins, 1000-1,000,000.

3. **Monomers** are usually made of carbon and hydrogen atoms. Oxygen, halogen, and nitrogen atoms can be present. Typical monomers are ethylene, $CH_2 = CH_2$, vinyl chloride, $CH_2 = CHCl$, vinyl acetate, $CH_2 = CHO_2CCH_3$.

4. **Addition polymers** are formed by monomers just "adding on" to each other without loss of atoms. All monomers forming addition polymers, have C=C bonds (carbon-carbon double bonds). These double bonds open up as the monomers join in chains 1000 to 10,000 units long. For example, ethylene molecules can join together (under the proper conditions of temperature, pressure, and catalyst) to form polyethylene:

   (See structures of monomers and addition polymers in *Appendix*.)

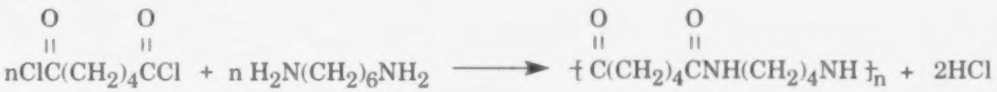
   $$nCH_2{=}CH_2 \xrightarrow{O_2\text{, heat, pressure}} \left(CH_2{-}CH_2\right)_n$$

5. Usually an **initiator** is used to start this polymerization reaction. The initiator is a compound that can form a free radical (a substance that contains an unpaired electron). When free radicals react with a double bond, a chain reaction occurs, which could theoretically continue as long as there is monomer material present. In the equation below, **RO•** is a generalized symbol for a free radical initiator.

$$CH_2{=}CH_2 + RO\bullet \longrightarrow RO{-}CH_2{-}CH_2\bullet \xrightarrow{CH_2=CH_2} RO{-}CH_2{-}CH_2{-}CH_2{-}CH_2\bullet$$

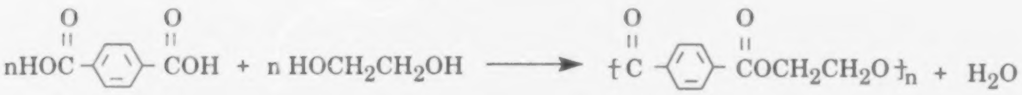
Other addition polymers include polymethyl methacrylate (Lucite™ or Plexiglass™), Teflon™, polyvinyl alcohol, polystyrene, polyvinyl chloride (PVC), polyacrylonitrile (Orlon™), and natural rubber (see *Appendix*).

6. **Condensation polymers** are formed when the monomers react with each other so that a small molecule, such as H_2O or HCl, is eliminated as a product each time a monomer reacts with another monomer. The formation of Nylon and Dacron are shown.



$$n\,\mathrm{Cl\overset{\overset{O}{\|}}{C}(CH_2)_4\overset{\overset{O}{\|}}{C}Cl} + n\,\mathrm{H_2N(CH_2)_6NH_2} \longrightarrow \left[\mathrm{\overset{\overset{O}{\|}}{C}(CH_2)_4\overset{\overset{O}{\|}}{C}NH(CH_2)_4NH}\right]_n + 2\mathrm{HCl}$$

Nylon



Dacron

There must be at least two functional groups on each reacting monomer in order for the reaction to continue to form a chain. After each elimination, reacting groups are available for additional condensations. Other examples of condensation polymers are polyurethane foam, epoxy glue, peptides and proteins (see table in *Appendix*).

7. Polymers are produced as long chains that sometimes have connections between them. If the chains are *not* connected to one another, the polymer typically can be melted and reshaped repeatedly. These polymers are called **thermoplastics**. Examples are polyethylene, nylon, polyester, polyvinyl chloride, and polystyrene. If the long chains are joined together with cross-links, the material becomes extremely rigid; once formed, these polymers do not soften or melt when heated. These polymers are **thermosets**. Examples are the plastics used to manufacture bowling balls, handles on kitchen utensils, Formica™ table tops, and some auto bodies.

Pattern Recognition

"Polymers are everywhere." Recognize the importance of polymers to chemical industry (see *Appendix*—A List of Common Polymer Products).

Problem Solving

1. Have students write equations for formation of polymers using unfamiliar monomers (see *Appendix*—Polymers from Monomers).
2. Ask students to determine the number and kind of monomer units from the structural'formula of a polymer.

Decision Making

1. List five polymer items used each week that could be made from other materials. Would you be willing to give up these polymer items for a substitute? Would the substitute last longer, be cheaper, be easier to dispose of than polymers? Consider these and other aspects before deciding.
2. Brainstorm ways to persuade your community to recycle plastics. What incentives would be necessary? What collection methods could be used? Propose a stepwise plan to carry out your ideas. Ask a grocery store to help recycle bags. Meet with a local fast food restaurant representative to discuss its plans for recycling.

HISTORY: ON THE HUMAN SIDE

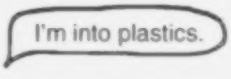
The history of polymers provides some excellent examples of serendipity. When Goodyear allowed a mixture of rubber and sulfur to touch a hot stove, he was surprised that it did not melt, but charred. Vulcanization was invented. The synthetic rubbers Neoprene and Thiocol were discovered accidentally. The usefulness of nylon fibers was discovered because chemists were trying to see how far they could stretch a ball of this new synthetic substance.

In his search for a nontoxic refrigerant, R. J. Plunkett of duPont opened a cylinder of tetrafluoroethylene, but no gas came out. Because the weight of the tank showed there was something inside, he sawed the tank open and found a white powder with inert properties. He discovered that the tank valve was the material that catalyzed polymerization of the gas to form the useful product Teflon™. It should be noted, however, that these serendipitous discoveries were made by scientists who were interested in these compounds and could recognize the conditions under which they were produced. It is true that "chance favors the prepared mind."

Wallace Hume Carothers (1896-1937) was the person primarily responsible for the development of nylon. This Iowa youth was fascinated by tools and mechanical devices. In 1915, he entered Tarkio College (Missouri) and excelled in chemistry, and became a chemistry instructor before he graduated. Eventually Carothers attended the University of Illinois where he completed his Ph.D. in organic chemistry in 1924 and was appointed to the faculty. In 1926, he moved to Harvard University but left two years later to join DuPont and participate in a new program of fundamental research. At DuPont, Carothers headed the organic chemistry division and played a prominent role in laying the foundations of polymer chemistry. By the age of 33, he had become world-famous and was the first industrial chemist to be elected to the National Academy of Sciences.

Carothers was a lover of poetry and classical music. Unfortunately, he suffered from severe bouts of depression that finally led to his suicide in a Philadelphia hotel at the age of 41. Despite his early death, Carothers' accomplishments were many, and he is considered one of America's greatest chemists.

HUMOR: ON THE FUN SIDE


I'm into plastics.

1. **CHEMTOON** $CH_2{=}CH_2$

   CHEM 13 NEWS, February 1985, p. 5

2. **Message on a bumper sticker**: Polymer chemists do it in chains!

3. Word Search (see *Appendix* for master copy)

~~C~~	Q	L	~~N~~	N	H	~~I~~	I	S	H	M	R	D	F	G	B	H
~~I~~	K	Y	~~O~~	M	R	~~N~~	B	P	M	T	L	O	N	P	Q	E
~~T~~	S	K	~~B~~	L	J	~~I~~	H	B	E	H	Q	E	D	R	H	C
~~S~~	R	A	~~R~~	T	I	~~T~~	T	Q	G	G	J	A	E	V	L	C
~~A~~	S	D	~~A~~	~~D~~	~~D~~	~~I~~	~~T~~	~~I~~	~~O~~	~~N~~	~~Y~~	Z	G	I	Y	F
~~L~~	Y	R	~~C~~	V	I	~~A~~	T	I	F	J	~~L~~	X	C	X	B	L
~~P~~	R	Q	Z	Y	F	~~T~~	~~S~~	~~T~~	~~E~~	~~S~~	~~O~~	~~M~~	~~R~~	~~E~~	~~H~~	~~T~~
~~O~~	U	F	W	~~P~~	~~R~~	~~O~~	~~T~~	~~E~~	~~I~~	~~N~~	~~P~~	R	~~E~~	F	W	A
~~M~~	~~O~~	~~N~~	~~O~~	~~M~~	~~E~~	~~R~~	Q	K	P	Q	Q	H	~~M~~	R	V	F
~~R~~	K	W	S	Q	O	J	O	W	P	X	E	G	~~Y~~	P	Y	S
~~E~~	R	T	M	G	A	J	A	F	A	Q	J	I	~~L~~	N	O	F
~~H~~	D	N	~~N~~	~~O~~	~~I~~	~~T~~	~~A~~	~~S~~	~~N~~	~~E~~	~~D~~	~~N~~	~~O~~	~~C~~	K	L
~~T~~	W	S	T	D	T	X	P	E	M	A	I	M	~~P~~	E	B	U

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Giant molecule formed from repeating units.
2. Individual unit of which polymers are made.
3. Compound used to start a polymerization reaction.
4. Type of polymer where repeating units "add on" to each other with no loss of atoms.
5. Type of polymer formed with the loss of a small molecule as monomers react.
6. Polymers that do not soften or melt when heated.
7. Polymer that can be melted and reshaped repeatedly.
8. Backbone element of most polymers.
9. Root word meaning "many."
10. Biological condensation polymer where amino acids are the monomers.

Answers: 1. POLYMER 2. MONOMER 3. INITIATOR 4. ADDITION 5. CONDENSATION 6. THERMOSETS 7. THERMOPLASTIC 8. CARBON 9. POLY 10. PROTEIN

4. See cartoons at end of module.

MEDIA

1. The World of Chemistry videotape "Number 22: Polymers." World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.
2. The Plastic Packaging Division of Mobil Chemical Company, 100 North Street, Canandaigua, NY 14424-1089 has two video cassettes on recycling, "Do the Right Thing: RECYCLE" and an interview with William Rathje by Connie Chung.
3. Philip Dail, 1416 Ewing Drive, Garner, NC 27529 has written a computer program on polymers that he will copy and send on request.
4. "Passing Through," informative and entertaining video introduction to plastics chemistry and technology. 27 min. Provided by National Plastics Center & Museum through Newtel Associates, 172 Lincoln Street, Worcester, MA 01605; (508) 791-8403.
5. Software published by *JCE:* Software, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   Viscosity of Polymer Solutions, by Gary L. Bertrand. Vol. V A, No. 2, for the Apple II computer.

6. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).
   a. For the Apple II computer running on ProDOS: AR 701
   b. For the Apple II computer: AP 702, AP 705, AP 209
   c. For IBM PCs and PC-compatibles: PC 3601, PC 4901
7. Videodiscs published by *JCE:* Software, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
   a. "From an Amino Acid to a Peptide Chain," "The Alpha Helix," "Hair and Silk," and "DNA Structure, Synthesis of Messenger RNA, Protein Synthesis," four chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc II (double sided, 60 min.), Special Issue 4.
   b. "Polymers," "Carbohydrates," and "Proteins," three chapters on the videodisc *Demonstrations in Organic Chemistry* (double sided, 60 min.), Special Issue 6.

Links/Connections

WITHIN CHEMISTRY

1. **Organic Chemistry**. Addition polymers are formed from monomers that are alkenes. Various condensation polymers are formed from aromatic compounds, acid chloride, carboxylic acids and amines (see *Addition and Condensation Polymers Transparency*).
2. **Biochemistry**. Many molecules important to living systems are polymers; for example, proteins (insulin is a polymer of amino acids), some carbohydrates (glycogen—the storage form of glucose), and the nucleic acids (DNA and RNA, the chemicals of heredity, are polymers of deoxyribose and ribose, respectively, and phosphoric acid).
3. **Catalysts**. Various conditions (temperature, pressure, catalyst) can be used to cause a monomer to polymerize. The catalyst can be one that generates positively charged intermediates—carbocations. A proton-donor can convert ethylene into polyethylene, as shown:

$$H^+ + CH_2{=}CH_2 \longrightarrow CH_3{-}CH_2^+ \xrightarrow{CH_2{=}CH_2} CH_3{-}CH_2{-}CH_2{-}CH_2^+ \quad \textit{etc.}$$

4. **Free Radicals**. Polymerization of alkenes can occur by a free-radical path. The reaction is initiated by O_2 or a peroxide (ROOR), as shown:

$$RO{-}OR \longrightarrow 2RO\bullet$$

$$RO\bullet + CH_2{=}CH_2 \longrightarrow RO{-}CH_2{-}CH_2\bullet \xrightarrow{CH_2{=}CH_2} RO{-}CH_2{-}CH_2{-}CH_2{-}CH_2\bullet \quad \textit{etc.}$$

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

1. **Use of Polymers in Medicine**. Dacron (a polyester of exceptional strength) fabrics have been used in surgery to replace or repair blood vessel segments. Nylon, a polyamide, is used in specialized tubing, and as velour for surfaces contacted by blood. Nylon sutures were the first synthetic sutures.
2. **Use of Polymers by Engineers**. Sound engineers store sound on compact discs, long-play records, and magnetic-coated audio- and videotapes that are made of polymers.
3. **Genetic Engineering**. Genetic engineering—controlled production of genetic changes—has made revolutionary advances in agriculture, medicine, and pharmaceutical manufacture. DNA, which carries the genetic information of heredity, is a polymer with repeating sugar and phosphoric acid monomers. Each sugar carries a base and each set of three consecutive bases will code ultimately for an amino acid in a protein (another polymer consisting of amino acid monomers). Scientists can read and synthesize specific genetic messages and have discovered special enzymes that split double-stranded DNA at specific locations. A new gene (with appropriate base triplets) could then be added or spliced to the original DNA and the resulting modified DNA repaired. The DNA—a new polymer—contains a new gene, and therefore, different hereditary information. Insulin, interferon, and human growth hormone have been made by this process.

4. **Physical Properties of Polymers**. Polymers are amorphous (noncrystalline) or crystalline. Noncrystalline polymers consist of a tangled mass of continuous chains or branched chains. Soft rubber is a noncrystalline polymer. Crystalline polymers have order and are composed of continuous chains held together by hydrogen bonds or other dipole-dipole attractions. Hard, moldable plastics are crystalline polymers.

5. **Environmental Studies**. Discarded plastic contributes to the garbage problem. Most plastic materials are not biodegradable. However, cornstarch is added to some plastics during manufacture. Bacteria eat the starch and the plastic breaks apart into small pieces. This method reduces the decomposition time from decades to several years.

   Some plastics—polyvinyl chloride and polyacrylonitrile—give off poisonous gases such as HCl and HCN when burned. Plastics are recyclable and are identified by codes (see *Plastic Container Code System for Plastic Bottles Transparency*). Such plastics can be separated from other plastics, remelted, and reused.

TO THE CONTEMPORARY WORLD

1. **Personal Use of Polymers** (see *A List of Common Polymer Products Transparency*). "Eco-Foam" is a new water-soluble starch which is now used as a packing material.
2. **Use of Polymers in Construction and Engineering**. Polymers are used to make numerous products that were formerly made of traditional materials like metals. Airplanes, boats, auto parts, chemical process equipment, appliance cabinets, electronic parts, business machine housings, and even caskets are now composed of polymers. Polymers are often lighter than the metal counterpart (less poundage), corrosion resistant (last longer), and parts can be consolidated so that a large complex part can be molded in one piece (rather than fabricated from many separate components, reducing assembly costs). In addition to the polymers, the materials used for these purposes are composites of reinforcing agents (for strength) and additives, such as flame retardants, flexibilizers, low-shrink agents, and pigments.
3. **Future of Polymers**. The raw materials for polymer manufacture come from petroleum and less from coal. These two fossil fuels are the main energy sources in the U.S. Consequently, polymers and energy are linked through these raw materials. Because the amount of these fossil fuels are limited, particularly petroleum, the items made of polymers derived from petroleum may have a limited lifetime.
4. **Field Trips**. Visit a solid waste facility and/or a recycling plant.
5. **Knowledgeable Individuals**. Invite environmentalists and representatives of industries and service organizations who recycle to speak. Invite a representative of your community government to speak on solid waste disposal. Invite an organic chemist or polymer chemist from academia or industry to speak on the chemistry of polymers.

References

Module developed by Jennifer Hubert, Jane A. Miller, and Marie C. Sherman, the Missouri team.

Alper, J. (1986). Polymers. *Chem Matters, 4*(2), 4.

History and current use and development of polymers. Especially focuses on synthetic ivory.

Butler, S., and Molote, S. (1981). Textile chemistry for the artist. *Journal of Chemical Education, 58*(4), 295.

Gives chemistry of textiles, especially fibers and color application, the knowledge of which can be valuable for artists, designers, and weavers.

Carragher, C. E., Jr., Hess, G., and Sperling, L. H. (1987). Polymer nomenclature—or what's in a name? *Journal of Chemical Education, 64*(1), 36.

Concentrates on naming of linear organic polymers. Four major types of names are also discussed: common, source-based, characteristic group, and structure based.

Cassasa, E. Z., Sarquis, A. M., and Van Dyke, C. H. (1986). The gelation of polyvinyl alcohol with borax. *Journal of Chemical Education, 63*(1), 57.

A novel class participation experiment involving the preparation and properties of a "slime."

Chiles, J. R. (1985, November). On land, at sea, and in the air, those polymer invaders are here. *Smithsonian*, 77.

Discusses the pervasiveness of plastics in every facet of our lives.

Cleary, J. (1986). Diapers and polymers. *Journal of Chemical Education, 63*(5), 422.

Illustrates how several polymers have been incorporated into the design of this "essential" product.

Cote, G. (1986). Polysaccharides. *Chem Matters, 4*(2), 12.

Different polysaccharides in food products. Focuses on seaweed polymers.

Deanin, R. D. (1987). The chemistry of plastics. *Journal of Chemical Education, 64*(1), 45.

Chemistry, development, and uses of plastic polymers.

Elliott, E. (1986). *Polymers and people.* Philadephia, PA: Beckman Center for the History of Chemistry.

Prepared to accompany a travelling exhibit. For more information, contact: Chemical Heritage Foundation Publications, 215 South 34th Street, Philadelphia, PA 19104-6310.

Ihde, J. (1990). Fun audience participation polymerization demonstration. *Journal of Chemical Education, 67*(3), 264.

Several student participation demonstrations illustrating addition polymerization that can be used at primary, secondary, and tertiary levels.

Kirshenbaum, G. S. (1987). Polymer literature for classrooms. *Journal of Chemical Education, 64*(1), 47.

Gives names and addresses of societies and associations with information on different types of polymers.

Kolb, K. E., and Kolb, D. K. (1988). Classroom analogy for addition polymerization. *Journal of College Science Teaching, 17*, 230.

This exercise helps students understand the basic concept of addition polymerization of ethylene-type monomers.

Mark, H. F. (1966). *Life science library: giant molecules.* New York: Time.

The story of polymers, the centuries of progress in chemistry that led to their creation, and the growth of research and production in this area.

Marsella, G. (1986). Silly putty. *Chem Matters, 4*(2), 15.

History, development, and chemistry of polymer now called "silly putty."

Morris, P. J. T. (1986). *Polymer pioneers.* Philadelphia, PA: Beckman Center for the History of Chemistry.

Historical overview of polymer science and technology.

Polymer Chemistry. (1986). NSTA Publication.

Rathje, W. L. (1989, December). Rubbish! *The Atlantic,* 99.

An archeologist who excavates modern landfills argues that much of the conventional wisdom about solid waste in America may be misleading or downright wrong.

Reshaping Our Lives: Advanced Materials. (1989). *National Geographic, 1176*, 746.

Man's increasing ability to manufacture new materials, plastics, ceramics, composites, and alloys, by manipulating the atoms of existing materials.

Sarquis, A. M. (1986). Dramatization of polymeric bonding using slime. *Journal of Chemical Education, 63*(1), 60.

Helps students visualize the type of reaction involved in polymerization using PVA, "slime." Even first graders can understand it.

Scheinberg, S. P. (1986). Tyvek. *Chem Matters, 4*(2), 8.

History, development and chemistry of plastic polymer in many products called Tyvek.

Stahl, G. A. (1981). *Polymer science overview: A tribute to Herman F. Mark.* Washington, DC: ACS Symposium Series.

A short history of polymer science.

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Three experiments show the properties of polymers; presented in the form of a party to generate student interest.

USA Today. (1990, June 20). On artificial fibers.

Wood, C. J. (1987, October). Dissolving plastic. *Chem Matters, 5*(3), 12-15.

Chemistry and uses of a water soluble plastic polymer, polyvinyl alcohol.

The following industry groups are involved in the solid waste issue:

Center for Plastics Recycling Research
Rutgers, The State University of New Jersey
Busch Campus, Bldg. 3529
Piscataway, NJ 08855
(201)932-3683

Council for Solid Waste Solutions
1275 K Street, NW, Suite 400
Washington, DC 20005
(202)371-5319

Flexible Packaging Association
1090 Vermont Avenue, NW
Washington, DC 20005
(202)842-3880

Keep America Beautiful, Inc.
Mill River Plaza
9 West Broad Street
Stamford, CT 06902
(203)323-8987

National Recycling Coalition
P.O. Box 80729
Lincoln, NE 68501-0729
(402)475-3637

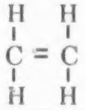
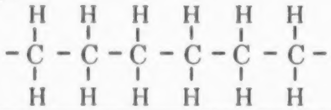
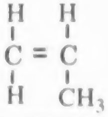
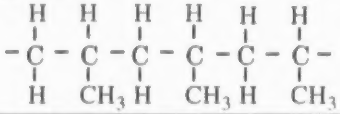
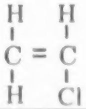
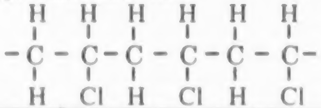
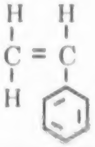
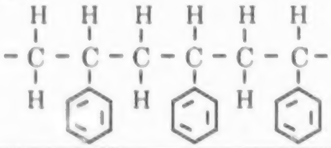
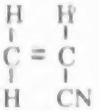
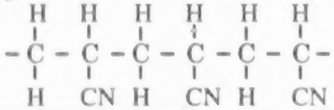
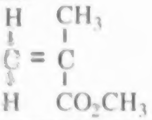
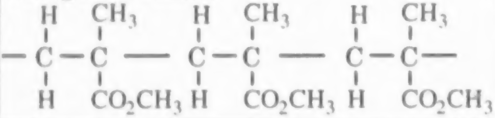
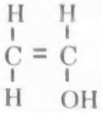
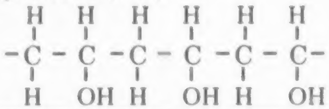
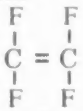
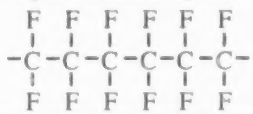
National Solid Wastes Management Association
1730 Rhode Island Avenue, NW, Suite 1000
Washington, DC 20036
(202)659-4613

National Plastics Center and Museum
P.O. Box 639
Leominster, MA 01453
(508) 537-9529

Appendix

- **Transparency Masters**
  1. Polymers from Monomers
  2. Addition and Condensation Polymers
  3. From Crude Oil to Plastics, I
  4. From Crude Oil to Plastics, II
  5. Plastic Container Code System for Plastic Bottles
  6. A List of Common Polymer Products
  7. Word Search
- **Humor**

Polymers from Monomers

Monomer	Polymer
Ethylene  $\begin{array}{ccc} H & & H \\ \vert & & \vert \\ C & = & C \\ \vert & & \vert \\ H & & H \end{array}$	**Polyethylene**  $\begin{array}{ccccccccccccc} & H & & H & & H & & H & & H & & H & \\ & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & \\ - & C & - & C & - & C & - & C & - & C & - & C & - \\ & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & \\ & H & & H & & H & & H & & H & & H & \end{array}$
Propylene  $\begin{array}{ccc} H & & H \\ \vert & & \vert \\ C & = & C \\ \vert & & \vert \\ H & & CH_3 \end{array}$	**Polypropylene**  $\begin{array}{ccccccccccccc} & H & & H & & H & & H & & H & & H & \\ & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & \\ - & C & - & C & - & C & - & C & - & C & - & C & - \\ & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & \\ & H & & CH_3 & & H & & CH_3 & & H & & CH_3 & \end{array}$
Vinyl Chloride  $\begin{array}{ccc} H & & H \\ \vert & & \vert \\ C & = & C \\ \vert & & \vert \\ H & & Cl \end{array}$	**Polyvinyl Chloride (PVC)**  $\begin{array}{ccccccccccccc} & H & & H & & H & & H & & H & & H & \\ & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & \\ - & C & - & C & - & C & - & C & - & C & - & C & - \\ & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & \\ & H & & Cl & & H & & Cl & & H & & Cl & \end{array}$
Styrene 	**Polystyrene** 
Acrylonitrile  $\begin{array}{ccc} H & & H \\ \vert & & \vert \\ C & = & C \\ \vert & & \vert \\ H & & CN \end{array}$	**Polyacrylonitrile**  $\begin{array}{ccccccccccccc} & H & & H & & H & & H & & H & & H & \\ & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & \\ - & C & - & C & - & C & - & C & - & C & - & C & - \\ & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & \\ & H & & CN & & H & & CN & & H & & CN & \end{array}$
Methyl Methacrylate  $\begin{array}{ccc} H & & CH_3 \\ \vert & & \vert \\ C & = & C \\ \vert & & \vert \\ H & & CO_2CH_3 \end{array}$	**Plexiglas™ or Lucite**  $\begin{array}{ccccccccccccc} & H & & CH_3 & & H & & CH_3 & & H & & CH_3 & \\ & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & \\ - & C & - & C & — & C & - & C & — & C & - & C & — \\ & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & \\ & H & & CO_2CH_3 & & H & & CO_2CH_3 & & H & & CO_2CH_3 & \end{array}$
Vinyl Alcohol  $\begin{array}{ccc} H & & H \\ \vert & & \vert \\ C & = & C \\ \vert & & \vert \\ H & & OH \end{array}$	**Polyvinyl Alcohol (PVA)**  $\begin{array}{ccccccccccccc} & H & & H & & H & & H & & H & & H & \\ & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & \\ - & C & - & C & - & C & - & C & - & C & - & C & - \\ & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & \\ & H & & OH & & H & & OH & & H & & OH & \end{array}$
Tetrafluoroethylene  $\begin{array}{ccc} F & & F \\ \vert & & \vert \\ C & = & C \\ \vert & & \vert \\ F & & F \end{array}$	**Teflon (Polytetrafluoroethylene)**  $\begin{array}{ccccccccccccc} & F & & F & & F & & F & & F & & F & \\ & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & \\ - & C & - & C & - & C & - & C & - & C & - & C & - \\ & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & \\ & F & & F & & F & & F & & F & & F & \end{array}$

Addition and Condensation Polymers

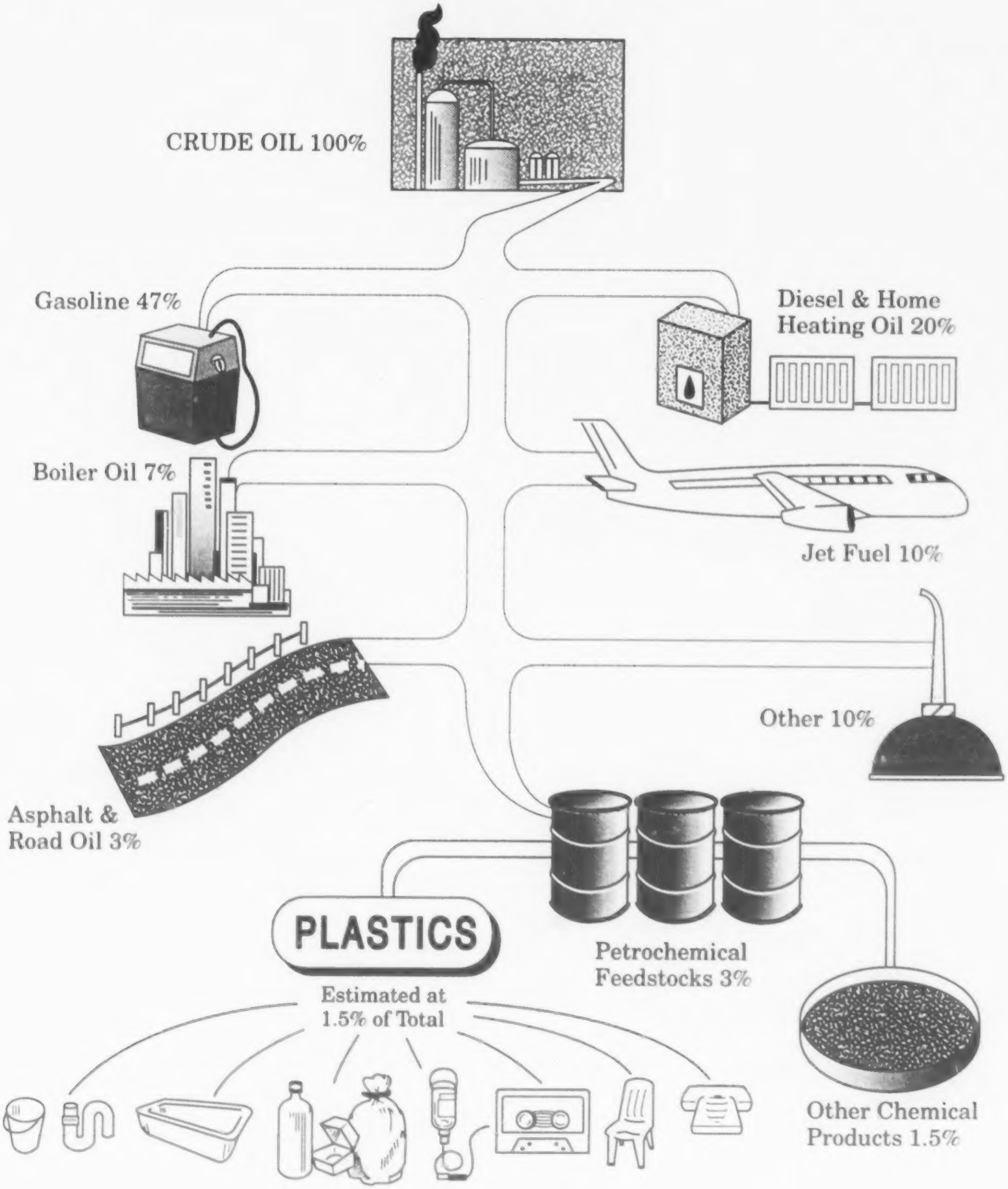
Addition Polymers

Monomer	Polymer	Uses
Ethylene	Polyethylene	Films, coatings, bottles, toys
Propylene	Polypropylene	Fibers, films, bottles, lab equipment
Vinyl chloride	Polyvinyl chloride (PVC)	Credit cards, phonograph records, floor tiles
Styrene	Polystyrene	Insulation, packing materials, coffee cups
Acrylonitrile	Polyacrylonitrile	Orlon, Acrilan, rug fibers
Vinyl acetate	Polyvinyl acetate	Latex paints
Methyl methacrylate	Polymethyl methacrylate	Glass substitutes, jewelry
Vinyl alcohol	Polyvinyl alcohol (PVA)	Glues
Tetrafluoroethylene	Polytetrafluoroethylene	Heat-resistant coatings

Condensation Polymers

Monomer	Polymer	Uses
Adipic acid, hexamethylene diamine	Nylon	Fibers, bearings
Orthophthalic acid, glycerine	Alkyd resins	Paints
Toluene diisocyanate, ethylene glycol	Polyurethane	Foam plastic, insulation
Dimethyl terephthalate, ethylene glycol	Dacron polyester Mylar	Fibers, plastic bottles

From Crude Oil to Plastics, I


CRUDE OIL 100%
Gasoline 47%
Diesel & Home Heating Oil 20%
Boiler Oil 7%
Jet Fuel 10%
Other 10%
Asphalt & Road Oil 3%
PLASTICS
Estimated at 1.5% of Total
Petrochemical Feedstocks 3%
Other Chemical Products 1.5%

Source: U.S. Energy Information Administration, *Petroleum Supply Annuals*

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From Crude Oil to Plastics, II

Most crude oil is turned into fuel for transportation. Less than 2 % is used as feedstock in making plastics. Here's a breakdown of all the ways the United States uses crude oil.

GASOLINE	Of all the crude oil refined for use in the United States, almost half (47%) becomes gasoline for automobiles, boats and other gasoline-driven motors.
JET FUEL	Airplanes consume 10%, in the form of jet fuel.
DIESEL FUEL AND HOME HEATING	Another 20% becomes distillate, two-thirds of which is diesel fuel for trucks, buses and other diesel engines, and one-third home heating oil.
BOILER OIL	Boiler oil, or residual fuel oil, which makes up 7% of crude oil consumption, is used on ships, in industrial boilers and in power plants to produce electricity.
ASPHALT AND ROAD OIL	Asphalt and road oil account for 3% of crude oil consumption.
OTHER	Some of the crude oil, about 10%, is used as non-energy feedstocks for manufacturing products such as lubricants, wax, coke for steel making, and naphthas that are used in the drycleaning process.
PETROCHEMICAL FEEDSTOCKS	Petrochemical feedstocks, products of the refining process, make up the remaining 3% of all crude oil consumption. Half are used to make PLASTICS (1.5% of the total) for thousands of items such as tableware, furniture, aircraft and automobile parts, luggage, surfboards, helmets, medical supplies and packaging. The remaining 1.5% is used to make products such as solvents, synthetic fibers for wearing apparel, synthetic rubber, paints and coatings.
IMPORTANT CONSIDERATION	Plastics bring about savings in energy use. For example, parts for cars and trucks are increasingly made of plastic to reduce their weight, and that means better fuel economy. Plastics packaging and other plastic products usually weigh less than their alternatives, and this saves fuel in shipping.

For more information contact (800) 333-0124, or write:
Mobil Chemical Company
Solid Waste Management Solutions
1159 Pittsford-Victor Road
Pittsford, NY 14534

Plastic Container Code System for Plastic Bottles

Code	Material
♳ PETE	Polyethylene terephthalate (PET)*
♴ HDPE	High-density polyethylene
♵ V	Vinyl/polyvinyl chloride (PVC)*
♶ LDPE	Low-density polyethylene
♷ PP	Polypropylene
♸ PS	Polystyrene
♹ Other	All other resins and layered multimaterial

* Stand alone bottle code is different from standard industry identification to avoid confusion with registered trademarks.

A List of Common Polymer Products

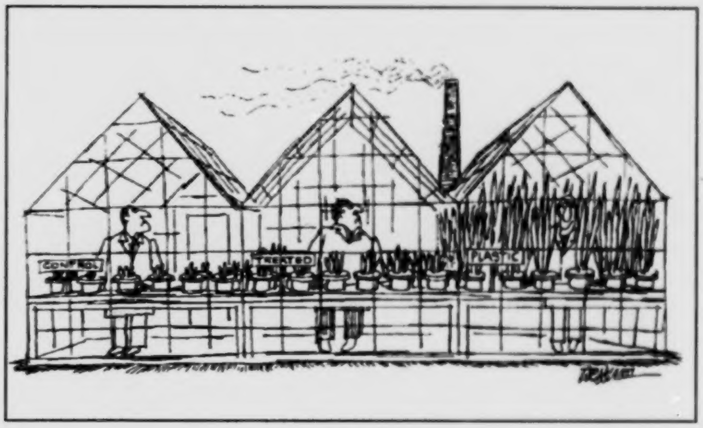
2-Liter pop bottle
ashtray
Astroturf
attache case
audio tape
awning
baby bottle
baby rattle
balloon
band-aid
basketball
beverage box
bicycle and automobile paint
bicycle and automobile tire
bicycle hand grip
boat
book bag
bubble gum
bubble pack
butane lighter
buttons
calorimeter cup (also for coffee)
camera
car battery
carpet
caulking
ceiling light cover
cellophane tape
chair seat
chapstick tube
checkers
chemistry classroom desk
clarinet
cleats
coaster
coffee mug
coffee stirrer
Colorforms
comb
compact
computer disc
computer
contact lenses
Corvette
credit card
curtains
cushion
dashboard
decorative fruit
dish pan
disposable diaper
disposable razor
dust brush
earphone
egg carton
electrical tape
epoxy glue
eraser
exercise mat
extension cord
false eyelashes
false teeth
fan belt
fastfood container
fishing bobber
fishing line
flag
flashcube
flea collar
floor "wax"
floor mat
floor tile
flower pot
flute
foam rubber pillow
folding door
food storage container
food wrap
football
football helmet
football pad
fork and spoon
Frisbee
furniture polish
galoshes
garbage bag
gears
glasses
guitar string
Habitrail
hair curler
hair dryer
hair spray
hang glider
hearing aid
hockey puck
Honda CRX
house paint
ice chest
ice cube tray
index tabs
insulated foam cup
insulation
jewelry
junk food bag
knapsack
lawn chair
life raft
lifesaver
lipstick tube
Lucite sculpture
luggage
lunch box
lunch tray
mannequin
margarine tub
measuring tape
metallic balloon
microfilm
microwave cookware
milk jug
model car
model plane
mop
motor oil bottle
motorcycle helmet
movie film
Nerf ball
overhead projector
pacifier
paint brush
pantyhose
parachute
particle board
patio screen
pencil case
pens
photographic film
photograph
piano
picture frame
pillow case
ping-pong ball
plastic flower
plastic dishes
plastic glasses and cups
plastic knives
playing cards
plywood
Pontiac Fiero
portable radio
pot handle
protractor
puppet
racquetball
raincoat
recorder
records
reflector
refrigerator
rope
rubber band
rubber duckie
rubber glove
rubber soles on shoes
rubber tubing
ruler
safety glass
safety glasses
sail
sailboat
sandals
sandwich bag
school desk
seat cushion
shampoo bottle
shoe box
shoe polish
shoe string tip
shoe string
shower door
shrink wrap
shuttlecock
Silly Putty
Silly String
skateboard wheels
slide
Slime
sneakers
snorkel
sofa
sousaphone
sponge
store sign
street sign
sun visor
sunglasses
Super Glue
supermarket meat tray
surfboard
sweater
sweater box
swim fins
synthetic fabric for clothing
table top
tackle box
tape recorder
Teflon cookware
telephone
tennis ball
tennis racquet and string
tent
tent peg
test-tube brush
test-tube rack
thermos bottle
thread
thread spool
toilet seat
toothbrush
toothpaste tube
transparencies
trash can
trophy
Tygon tubing
typewriter case
typewriter ribbon
umbrella
Velcro
venetian blinds
videotape
vinyl siding
vinyl top
vinyl wall covering
vitamin capsule
volleyball and net
waders
wallet
watch face
water pipe
Weed Whacker
welcome mat
wet suit
whistle
wig
windbreaker
windshield
windshield wiper
yarn
zipper

Word Search

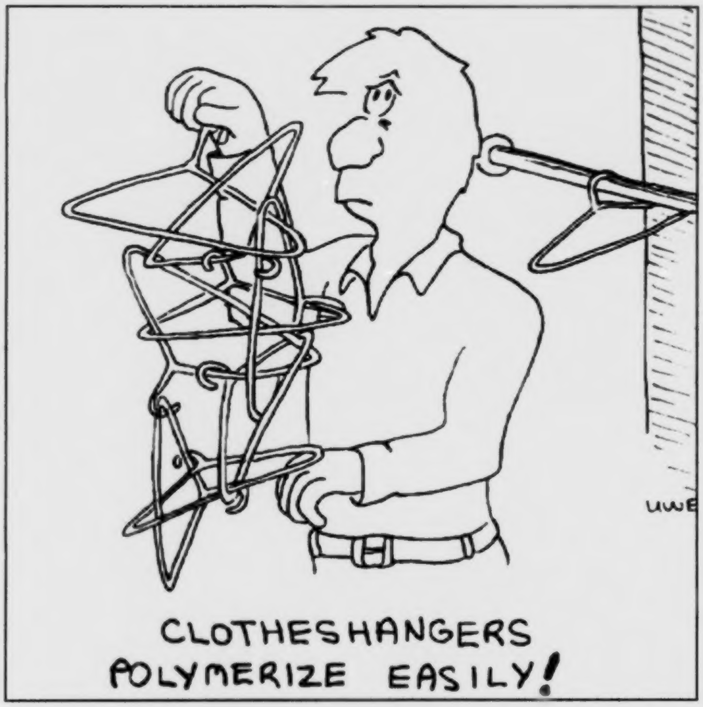
```
C Q L N N H I I S H M R D F G B H
I K Y O M R N B P M T L O N P Q E
T S K B L J I H B E H Q E D R H C
S R A R T I T T Q G G J A E V L C
A S D A D D I T I O N Y Z G I Y F
L Y R C V I A T I F J L X C X B L
P R Q Z Y F T S T E S O M R E H T
O U F W P R O T E I N P R E F W A
M O N O M E R Q K P Q Q H M R V F
R K W S Q O J O W P X E G Y P Y S
E R T M G A J A F A Q J I L N O F
H D N N O I T A S N E D N O C K L
T W S T D T X P E M A I M P E B U
```

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

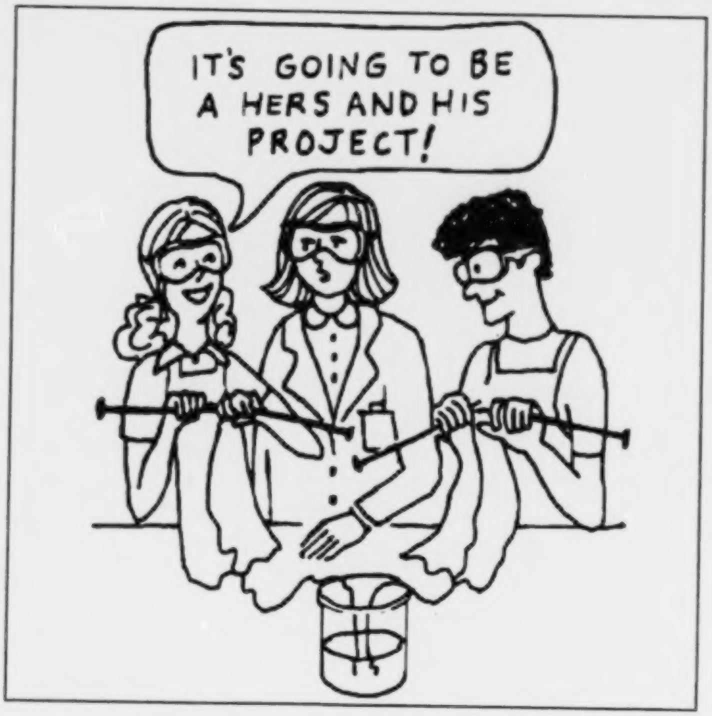
1. Giant molecule formed from repeating units.
2. Individual unit of which polymers are made.
3. Compound used to start a polymerization reaction.
4. Type of polymer where repeating units "add on" to each other with no loss of atoms.
5. Type of polymer formed with the loss of a small molecule as monomers react.
6. Polymers that do not soften or melt when heated.
7. Polymer that can be melted and reshaped repeatedly.
8. Backbone element of most polymers.
9. Root word meaning "many."
10. Biological condensation polymer where amino acids are the monomers.


CONTROL
TREATED
PLASTIC

CHEMTECH, November 1973, p. 688. © American Chemical Society. Reprinted with permission.

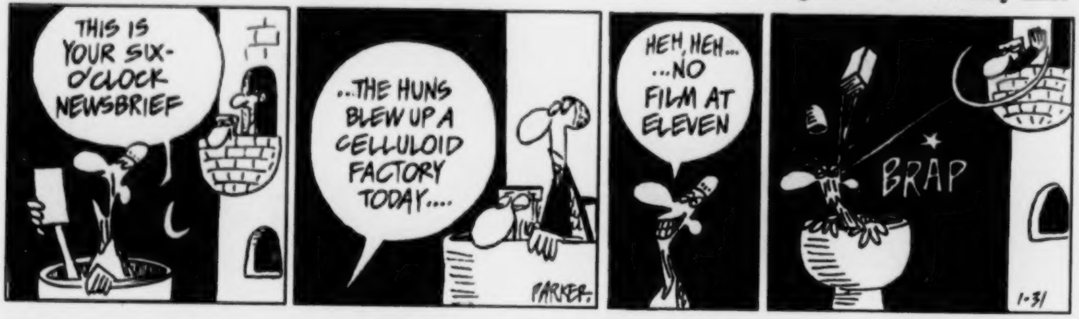

UWE
CLOTHESHANGERS POLYMERIZE EASILY!

CHEM 13 NEWS, September 1980, p. 4. Reprinted with permission.

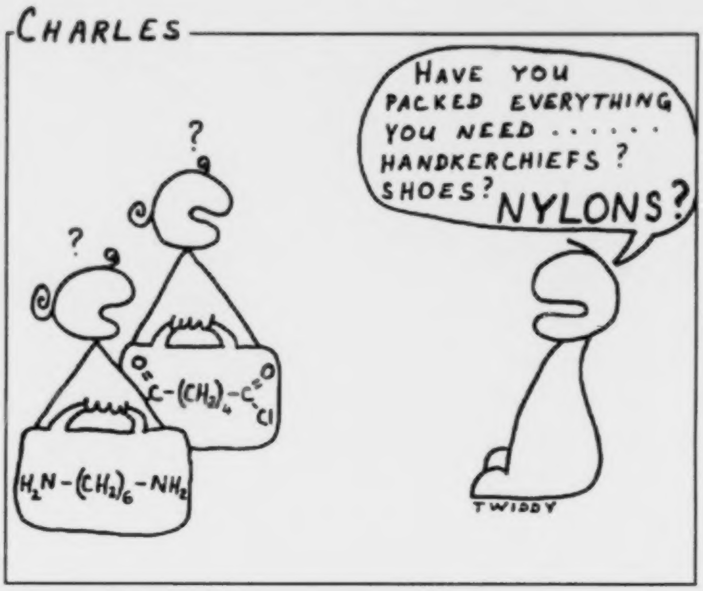

IT'S GOING TO BE A HERS AND HIS PROJECT!

Used by permission of Veritex Publishing Co.

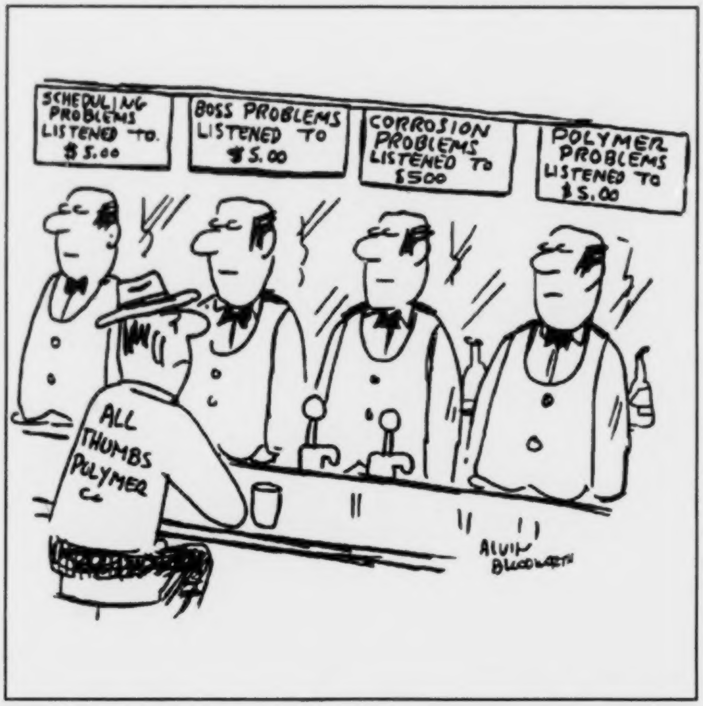
THE WIZARD OF ID **Brant parker and Johnny hart**


THIS IS YOUR SIX-O'CLOCK NEWSBRIEF
...THE HUNS BLEW UP A CELLULOID FACTORY TODAY....
PARKER
HEH, HEH... ...NO FILM AT ELEVEN
BRAP
1-31

By permission of Johnny Hart and Creators Syndicate, Inc.


CHARLES
HAVE YOU PACKED EVERYTHING YOU NEED ...... HANDKERCHIEFS ? SHOES? NYLONS?
?
?
$O=C-(CH_2)_4-C(=O)-Cl$
$H_2N-(CH_2)_6-NH_2$
TWIDDY

CHEM 13 NEWS, January 1985, p. 3. Reprinted with permission.


SCHEDULING PROBLEMS LISTENED TO. $5.00
BOSS PROBLEMS LISTENED TO $5.00
CORROSION PROBLEMS LISTENED TO $500
POLYMER PROBLEMS LISTENED TO $5.00
ALL THUMBS POLYMER Co
ALVIN BLOODWORTH

CHEMTECH, August 1972, p. 453. © American Chemical Society. Reprinted with permission.

Itsu-Butsi Plastics Company Limited?

CHEMTECH, March 1972, p. 135. © American Chemical Society. Reprinted with permission.

A SourceBook Module

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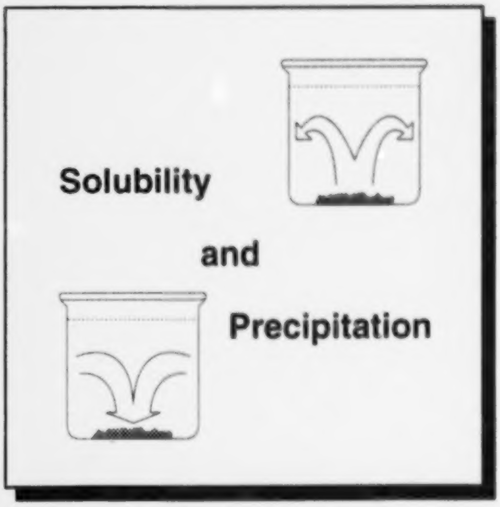
New Rochelle, NY 10805
Phone: (914) 654-5302
FAX: (914) 654-5387

SOLUBILITY AND PRECIPITATION



ChemSource

Instructional Resources for Preservice and
Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

Life on earth depends upon the fact that certain substances dissolve in other substances. We constantly deal with solutions. We enjoy cola drinks because dissolved carbon dioxide gives them an effervescence. Fish can live because oxygen gas is dissolved in lakes and oceans. We grow and develop because nutrients dissolved in our blood can reach cells and tissues.

The process of dissolving has always been fascinating. Early alchemists searched for the "universal solvent"—that mythical liquid that would dissolve everything. Perhaps it never occurred to them that, if they found it, it would be difficult to contain! Of course, there is no universal solvent. Water is perhaps the closest thing we have to a universal solvent; its truly unique properties make it one of our most important chemical substances.

We now know much more about the process of dissolving than did early chemists. We now understand that particles (ions and molecules) are held together within solids and liquids by certain attractive forces. Some of these are weak and some are relatively strong. For a solution to form, bonds must be broken and reformed, involving changes in energy. Solution formation is most likely to happen when the solute and solvent have certain similar chemical properties. We often generalize this fact by saying that "like dissolves like."

When a solute dissolves in a solvent in a closed system, the process continues until a certain amount of solute enters solution. After that point, particles return to the solute at the same rate they leave the solute and dissolve. Thus, dissolving is a dynamic equilibrium process. It has often been stated that "everything is somewhat soluble in everything else," and every substance has a characteristic solubility. When the solubility of a solid is exceeded, the substance comes out of solution in the form of a precipitate. The word *precipitate* comes from the Latin and means "cast down." Interestingly enough, chemists a hundred years ago knew little about ions and atoms and how they recombine in solution. When a precipitate formed in solution, they would say that a substance was "thrown down."

Some substances show very limited solubility in solvents like water. We can express this small solubility quantitatively as the "solubility product constant" (K_{sp}). This is a useful expression since it allows a chemist to predict if and when a precipitate may occur when solutions are mixed.

The solubility of a substance can be changed by altering the conditions that affect equilibrium. Temperature and pressure affect the solubility of gases in liquids. The solubility of solids in liquids is affected by temperature.

If we collect enough data regarding the solubility of certain substances, especially ionic compounds, we can formulate a set of rules or generalizations that help us predict whether or not certain compounds will form solutions.

PLACE IN THE CURRICULUM

Students can understand and appreciate the ideas of solubility and precipitation with even greater comprehension after they have studied the concepts of bonding and equilibrium as part of class discussion on the solution process. There are many important and practical applications of these concepts.

CENTRAL CONCEPTS

1. Certain substances have the ability to dissolve other substances.
2. The extent of solubility depends upon interparticle solute-solute, solvent-solvent, and solute-solvent attractions.
3. When the solute is only slightly soluble, the extent of solubility can be quantitatively expressed as the solubility product constant.
4. The solubility of a substance can be altered by the common ion effect, pH change, temperature change, *etc.*
5. Accumulated data allow one to predict the degree of solubility of many substances (solubility rules).

RELATED CONCEPTS

1. Bonding
2. Energy
3. Entropy
4. Polar and nonpolar molecules
5. Ionization/dissociation
6. Equilibrium
7. Kinetic molecular theory

RELATED SKILLS

1. Exponents
2. Constants
3. Logarithms
4. Working with small-scale volumes
5. Measuring volumes using a graduated cylinder or other volumetric apparatus.

PERFORMANCE OBJECTIVES

After completing their study of solubility and precipitation, students should be able to:

1. explain phrase "like dissolves like" as it relates to polar and nonpolar materials.
2. discuss the effect that entropy and enthalpy have on the dissolving process.
3. explain why a precipitate forms when solutions of two ionic compounds are mixed.
4. explain why gases become less soluble at higher temperatures, whereas most solids become more soluble.
5. predict the solubility of an ionic compound by applying generalizations (rules) determined by laboratory experiments.
6. give practical, personal, and everyday examples of situations that involve solubility and precipitation.
7. explain the relationship between equilibrium and solubility.
8. explain how solubilities of solids and gases are changed by changing temperature and pressure.
9. discuss effect of temperature on endothermic and exothermic solution processes.
10. describe how increase in disorder favors the solution process.
11. discuss energy required for bond breaking *vs.* energy released when new bonds are formed.
12. calculate the K_{sp} for an ionic compound if they are given the solubility of the compound.
13. calculate solubility of an ionic compound if they are given K_{sp} for that compound.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: What Dissolves and What Does Not: Establishing Some "Rules"

Introduction

In this laboratory activity you will work with an assortment of solutions of ionic substances, divided into five sets. Your teacher will assign specific sets of solutions for your work. You will compile your data with that of classmates to generate some 'rules' regarding solubilities of certain ionic compounds. Using very small quantities, you will mix all possible combinations of solutions in each set and observe any reactions that occur.

Purpose

When you mix certain solutions, you will find that some combinations produce precipitates. By comparing data with other students, you should find distinctive patterns developing as certain ions react with each other. From such class data, some generalizations regarding the solubilities of ionic compounds can be drawn. This activity will give you experience in (a) developing laboratory skills, (b) making and recording observations, (c) drawing conclusions and making generalizations from data, and (d) working as part of a team.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Although none of the solutions are toxic, place all waste solutions in a properly labeled container. Do not discard the waste in the sink.

Procedure

1. Prepare five grid sheets like the following one. (The numbers refer to numbered solutions in each set.)

Set No. _____	6	5	4	3	2
1					
2					
3					
4					
5					

2. Select a set of solutions for the laboratory activity. (Your teacher may select one for you.) Indicate the number of the solution set on the sheet. For each solution in the set, identify the ions present in the solution and record them on the sheet. For example, the sheet for Solution Set No. 1 would look like the one shown. Prepare two sheets for each set of solutions.

Set No. 1	K^+/Cl^- 6	Na^+/NO_3^- 5	Ag^+/NO_3^- 4	Na^+/I^- 3	Ba^{2+}/Cl^- 2
Ba^{2+}/NO_3^- 1					
Ba^{2+}/Cl^- 2					
Na^+/I^- 3					
Ag^+/NO_3^- 4					
Na^+/NO_3^- 5					

3. Place an acetate sheet, or a small glass plate, over one of the sheets. (Your teacher may prefer that you use cell well plates.)
4. Carefully add one drop of each solution, where it is indicated on the grid. When you have finished, you will have two drops of solution in each square, and you will have mixed all possible combinations of solutions in the set.
5. Note any precipitates that form, and record their colors on the second grid sheet that you prepared. Write a net ionic equation for each reaction that produced a precipitate.
6. Carefully rinse the acetate sheet, or glass plate, with a small amount of water into the waste container provided by your teacher. Wash your hands.
7. Repeat the procedure for the other solution sets assigned to you.
8. Thoroughly wash your hands before leaving the laboratory.
9. Compile your data with that of your classmates on the large data table on the chalkboard.

Data Analysis and Concept Development

1. Observe the patterns in precipitation. What conclusions can you draw regarding the solubility of various combinations of ions in solution?
2. Which of the following pairs of ions should produce a precipitate in water solution?

a. $Ag^+ + NO_3^-$

b. $Mg^{2+} + OH^-$

c. $K^+ + SO_4^{2-}$

d. $Al^{3+} + Na^+$

e. $Ag^+ + CO_3^{2-}$

f. $NH_4^+ + Cl^-$

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: What Dissolves and What Does Not: Establishing Some "Rules"

Level

This activity may used with basic, general, and honors students.

Safety

These solutions include some heavy metal ions that are toxic in the environment. Although they are safe to use when handled properly, you must dispose of these ions carefully. A small-scale approach is suggested to reduce the volume of solutions required.

Materials

	Set I	Set II	Set III	Set IV	Set V
Solution 1:	$Ba(NO_3)_2$	Na_2SO_4	$FeCl_3$	$CoCl_2$	$BaCl_2$
Solution 2:	$BaCl_2$	$Al_2(SO_4)_3$	$Co(NO_3)_2$	$MgCl_2$	$Sr(NO_3)_2$
Solution 3:	NaI	$Sr(NO_3)_2$	$CoCl_2$	K_2SO_4	Na_2CO_3
Solution 4:	$AgNO_3$	$BaCl_2$	NaOH	NaOH	$Al_2(SO_4)_3$
Solution 5:	$NaNO_3$	$Ba(NO_3)_2$	KOH	$Ba(OH)_2$	$K(CH_3COO)$
Solution 6:	KCl	$AlCl_3$	$NaNO_3$	$MgSO_4$	$AgNO_3$

Advance Preparation

1. Prepare 10% solutions of the reagents by dissolving 10 g of each in 90 mL water. If all of the sample does not dissolve let it settle and use the clear solution.
2. Place the solutions in dropper bottles and arrange in sets. Prepare at least five complete set-ups. These solutions are stable and can be used many times. **Assign only one set to each student pair.** (Beral™ pipets in an audio cassette case are also a convenient arrangement.)
3. Acetate sheets, or small glass plates work well. Acetate sheets are inexpensive and eliminate the risk of glass cuts. You may wish to photocopy copies of the grid and give several to each student, or make transparencies of the grid. Small-scale cell well plates work nicely. Students can rinse them and dry with Q-tips. Grid sheets on colored paper enhance visibility of white precipitates.
4. You might find it convenient to provide a large chart on the chalkboard and have students record their data as they complete each part of the laboratory activity. The following chart is suggested as a possible format.

	Cl^-	CH_3COO^-	CO_3^{2-}	OH^-	NO_3^-	I^-	SO_4^{2-}
Ag^+							
Al^{3+}							
Ba^{2+}							
Co^{2+}							
Fe^{3+}							
K^+							
Mg^{2+}							
Na^+							
Sr^{2+}							

Place an (I) for Insoluble in the space where a precipitate was formed.
Place an (S) in spaces where there was no precipitate.

Generally, we consider anything insoluble if it has a solubility of less than 0.1 g per 100 mL. You may find it necessary to declare some compounds as "slightly soluble" (ss). Silver sulfate fits this category.

Pre-Laboratory Discussion

1. When students compile their data and compare their observations, it will become apparent that certain combinations of solutions produce precipitates. Further, it can be determined that only certain pairs of ions react to produce precipitates. When class data are gathered, some generalizations regarding the solubility of ionic compounds can be drawn. It is important that the general rules for solubility be developed as a result of student observations, and not presented as something to be "memorized".
2. This activity gives students experience in (1) developing laboratory skills, (2) making and recording observations, (3) drawing conclusions and making generalizations from data, and (4) working as part of a team.
3. Students should be introduced to dissociation, the separation of ions in aqueous solution. They should identify all ions present in a solution and write correct ionic symbols. They must understand that precipitates form only when an insoluble product results from the recombination of ions in solution. They should also be able to write net ionic equations for precipitation reactions, eliminating "spectator" ions.
4. You might select, as an example, the combination of silver nitrate and sodium chloride solutions. Do this as a demonstration—simply add one to the other; a dense white precipitate of silver chloride will form. Ask students to list all ions present in the combined solutions and eliminate those that would not precipitate. Show students that the precipitate could not be sodium nitrate by dissolving a small amount of this compound in water to show that it is very soluble. Stress that the "process of elimination" to determine that the precipitate must be silver chloride is based on all four potential salts in each reaction plus recording all six solution salts as soluble in the beginning of the activity.

Teacher-Student Interaction

Having students pool and compare their data provides an opportunity to discuss experimental error and the need to collect a large quantity of data before establishing conclusions. Students should be prepared to support their conclusions with careful observations and data collection. They should also realize that the rules, theories, and hypotheses that chemists use are the direct or indirect result of laboratory experimentation. This laboratory activity allows every student to contribute to forming the "rules" by which chemists try to describe and understand their surroundings.

Anticipated Student Results

Student results will be encouragingly consistent. If some question arises as to whether or not a precipitate formed, you can mix several milliliters of the solutions in question in front of the class and ask for their observations.

Answers to Data Analysis and Concept Development

1. The patterns match those found in general rules for water solubilities of ionic substances.
2. a. No precipitate
   b. Precipitate
   c. No precipitate
   d. No precipitate
   e. Precipitate
   f. No precipitate

Post-Laboratory Activities

1. After discussing student data and creating a set of generalizations regarding solubilities of certain compounds, call attention to the "solubility rules" in the textbook, or distribute a copy. Compare these two sets of "rules." Spot check student understanding by posing possible combinations of solutions. (What would happen if solutions of lead nitrate and potassium carbonate were mixed? What would the precipitate be?)
2. You may find that barium hydroxide and sodium hydroxide (Set 4) produce a precipitate due to dissolved carbon dioxide in the sodium hydroxide, forming barium carbonate. If this happens, explain to students that it is not due to the combination of ions from the original compounds.

Assessing Laboratory Learning

1. It is important that students understand what their laboratory activity has produced regarding knowledge of solubility rather than having students simply memorize solubility rules. Allow students to refer to the class-generated solubility chart and predict the solubility of various ionic compounds, or the precipitate formed when various ionic solutions are mixed.
2. Ask students to design sets of solutions to check the solubility of compounds not included on their charts.
3. Have students write net ionic equations, eliminating all spectator ions.

DEMONSTRATIONS

Demonstration 1: Solubility and Immiscibility

Description

Polar solvents dissolve polar and ionic solids, and nonpolar solvents generally dissolve nonpolar solids. This can be easily demonstrated and projected on an overhead projector.

Materials

Trichlorotrifluoroethane (TTE), 5 mL
Solid iodine, $I_2(s)$, several crystals
Solid copper(II) sulfate, $CuSO_4 \cdot 5\ H_2O(s)$, several crystals
Petri dishes
Overhead projector

Procedure

Place two Petri dishes on the stage of an overhead projector. Add enough water to each dish to cover the bottom. In each dish, carefully add about 2 mL TTE. (Notice that it forms a "pool".) Carefully drop a few crystals of iodine in the TTE pool of one Petri dish, and the same number in the water near the pool. What do you observe? Now, in the second dish add a few crystals of copper(II) sulfate in the TTE pool, and the same amount in the water near the pool. Stir each gently with a toothpick. What do you observe?

Reactions

TTE is a nonpolar solvent; it does not dissolve in polar water. Iodine is a nonpolar solid; it does not dissolve in polar water, but does dissolve in nonpolar TTE. Copper(II) sulfate is an ionic solid that dissolves in polar water, but not in nonpolar TTE.

Disposal

TTE is an acceptable and safe organic solvent. All of these materials can be disposed of according to directions given in the *SourceBook Safety* section.

Demonstration 2: Negative Coefficient of Solubility

Description

The solubility of most solids increases with an increase in temperature. Some compounds, however, show a decrease in solubility with increased temperature. Calcium acetate, $Ca(CH_3COO)_2$, is such a compound—it has a negative coefficient of solubility.

Materials

Calcium acetate solution, saturated (about 60 g/150 mL water)
Thermometer
Erlenmeyer flask, 250-mL
Hot plate or burner

Procedure

Place about 150 mL calcium acetate solution in a 250-mL flask and cool it overnight. Heat the flask gently and note that a precipitate begins to form as it is heated (at about 80 °C). Remove the flask from the heat and cool it by placing it in a stream of tap water. Notice that the precipitate goes back into solution as the temperature decreases.

Disposal

Collect waste chemicals in a special container and dispose of them according to the acceptable procedure for your school. (This solution can be stoppered and saved from year to year. You may need to readjust the amount of water as needed after long periods of standing.)

Demonstration 3: Producing a Blue Precipitate

Description

Students often get the impression that all precipitates are white, since this is what they usually see. This demonstration produces a nice blue precipitate.

Materials

Limewater (saturated calcium hydroxide solution), ~0.5 g in 200 mL H_2O
0.1 M Cobalt(II) nitrate solution [2.9 g Cobalt(II) nitrate hexahydrate, $Co(NO_3)_2 \cdot 6\ H_2O$, per 100 mL solution] in dropping bottle

Procedure

Place the limewater in a large beaker. Add a few drops of cobalt nitrate solution. Note the immediate formation of a blue precipitate.

Precautions

Cobalt nitrate should be handled with care. Wear gloves and dispose of the precipitate by drying and following directions in the *SourceBook Safety* section.

Reactions

Cobalt ion reacts with the basic limewater solution to produce cobalt hydroxide, a blue precipitate.

$$Co^{2+}(aq) + 2OH^-(aq) \rightleftharpoons Co(OH)_2(s)$$

Demonstration 4: Precipitation and Purification of Water

Description

Water that we use in our homes often comes from rivers, lakes, or ground deposits, and contains impurities such as bacteria, soil, and toxic substances that must be removed before we can use it. After water is allowed to stand in large tanks to allow large particles to fall to the bottom, it is treated to remove smaller suspended particles. This demonstration duplicates that treatment.

Materials

Lime solution (calcium hydroxide; about 3 tsp per liter water)
Alum solution (potassium aluminum sulfate, or aluminum sulfate, or ammonium aluminum sulfate; about 3 tsp per liter water)
Red litmus paper
Soil
Containers

Procedure

Add enough soil to a container of water to make a muddy solution. Divide this in half—one will be treated and one will serve as a control. Place a piece of red litmus paper in one container. Add lime solution, a dropperful at a time, stirring after each addition until the paper turns blue. Add a dropperful of alum solution to the container. What do you observe? Continue adding alum solution until a thick, white gelatinous precipitate forms. Stir the solution thoroughly and observe the two containers side-by-side for several minutes.

Reactions

The addition of lime solution, $Ca(OH)_2$, produced a basic solution. Aluminum ions reacted with this solution to produce aluminum hydroxide. This compound has a low solubility, so it forms a gelatinous precipitate that settles to the bottom, taking with it much of the suspended matter, leaving clear water.

$$Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$$

Disposal

This material can be flushed down the sink.

Counterintuitive Examples and Discrepant Events

1. Can you prepare a saturated solution of bromine in carbon tetrachloride? Ethyl alcohol in water? *[No, bromine, which is ordinarily a liquid, is infinitely soluble in carbon tetrachloride; thus it is impossible to form a saturated solution of these two liquids. Ethyl alcohol is also infinitely soluble in water.]*
2. When sugar dissolves in water, why is the resulting solution clear? Why has the sugar become "invisible?" *[You cannot see anything that is smaller than the wavelength of visible light (about 10^{-5} cm). When solid, visible sugar dissolves, it releases individual sugar molecules. These have a diameter of about 10^{-7} cm. Even when surrounded by water molecules, their size is less than the wavelength of visible light. Since lightwaves pass through them, they are, indeed, "invisible."]*
3. Why do the particles in a true solution never settle? *[Particles in solution have extremely low mass. A sugar molecule, for example, has a mass of about 5×10^{-22} g. Although even a particle this small is affected by gravity, it is constantly bombarded by water molecules. As it begins to fall, water molecules below hit it and tend to drive it upward, keeping it in solution. Particles large enough to overcome this bombardment will settle, and will not form a true solution.]*
4. Do all solids that lose their physical state in aqueous solution dissolve in water? *[No. Salts dissolve, but metals react with acids and then dissolve because salts are formed.]*

Analogies

1. Ions separating from compounds and recombining to form precipitates may be compared to couples (cations + anions) going to a "dance." During the dance, couples exchange partners to form "new combinations."
2. Increase in entropy during the dissolving process is somewhat like a small group of students going into a large auditorium, where they have a tendency to spread out, rather than bunch together.

Pictures in the Mind

The Dissolving Process

1. Draw a picture of the water molecule, showing its positive and negative "ends."
2. Draw a picture of a sodium chloride crystal, showing its positive and negative ions.
3. Draw a picture of several water molecules attracted to a negative chloride ion.
4. Draw a picture of several water molecules attracted to a positive sodium ion.

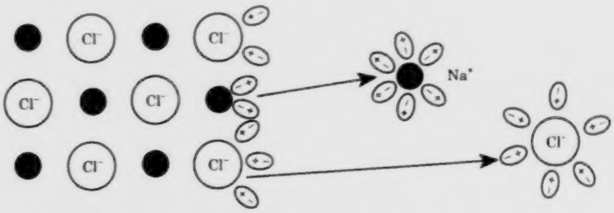


Figure 1. The dissolving process.

The Precipitation Process

1. Draw a picture of two different ions in a solution.
2. Draw a picture of a different solution with two different ions.
3. Draw a picture of all these ions in the two solutions mixed together.
4. Draw a picture to show how a precipitate might form from a new combination of these ions.

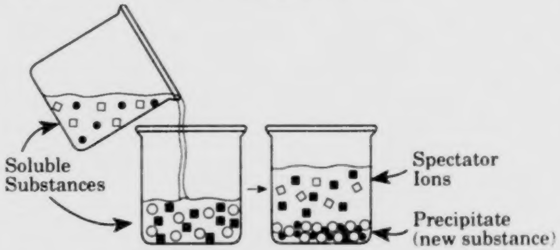


Figure 2. The precipitation process.

Language of Chemistry

Vocabulary

precipitate solid formed in solution by a reaction.

solubility amount of a substance that dissolves in a given quantity of solvent (for example water) at a given temperature.

solubility-product constant (K_{sp}) equilibrium constant for the dissolving of a slightly soluble (or nearly insoluble) ionic compound.

tincture an alcohol solution, as in "tincture of iodine."

TIPS FOR THE TEACHER

Major Factors in the Dissolving Process

$$\Delta H_{solute} + \Delta H_{solvent} + \Delta H_{mixing} = \Delta H_{solution}$$

If ΔH in "A" is GREATER than ΔH in "B," then

- $\Delta H_{solution}$ is Endothermic
- Solution gets colder as it forms
- Heating will cause more solid to dissolve
- Energy needed to break old attractive forces > Energy released by formation of new forces

If ΔH in "A" is LESS than ΔH in "B," then

- $\Delta H_{solution}$ is Exothermic
- Solution gets warmer as it forms
- Heating will decrease the amount of solid dissolved
- Energy needed to break old attractive forces < Energy released by formation of new forces

THE HEAT EFFECT (ENTHALPY)

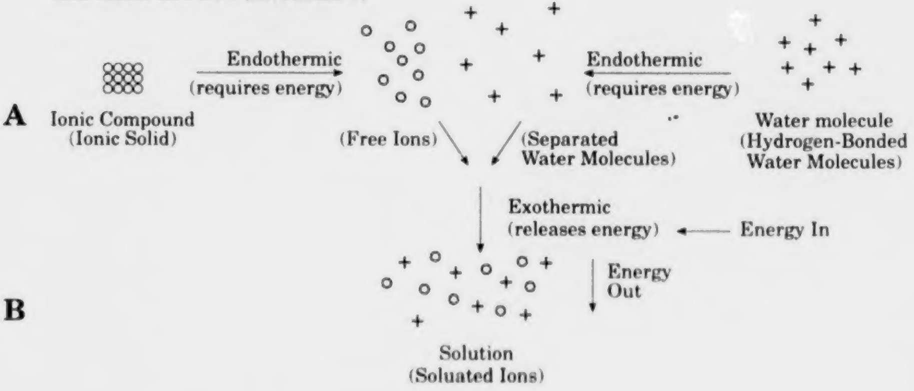

A
Ionic Compound
(Ionic Solid)
Endothermic
(requires energy)
(Free Ions)
(Separated
Water Molecules)
Endothermic
(requires energy)
Water molecule
(Hydrogen-Bonded
Water Molecules)
Exothermic
(releases energy)
Energy In
Energy
Out
B
Solution
(Soluated Ions)

Figure 3. The effect of heat (enthalpy).

THE TENDENCY TOWARD DISORDER EFFECT (ENTROPY)

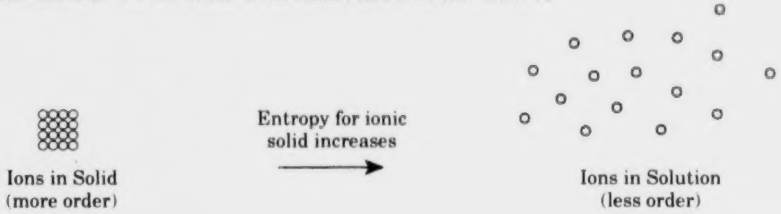

Ions in Solid
(more order)
Entropy for ionic
solid increases
Ions in Solution
(less order)

Figure 4. The effect of entropy.

There is a tendency for entropy to increase when solutions form.

If ΔS is positive, solution is favored.
ΔS is large enough, it can overcome ΔH effects and substance will go into solution.

These conditions for dissolving can also be represented by the following charts.

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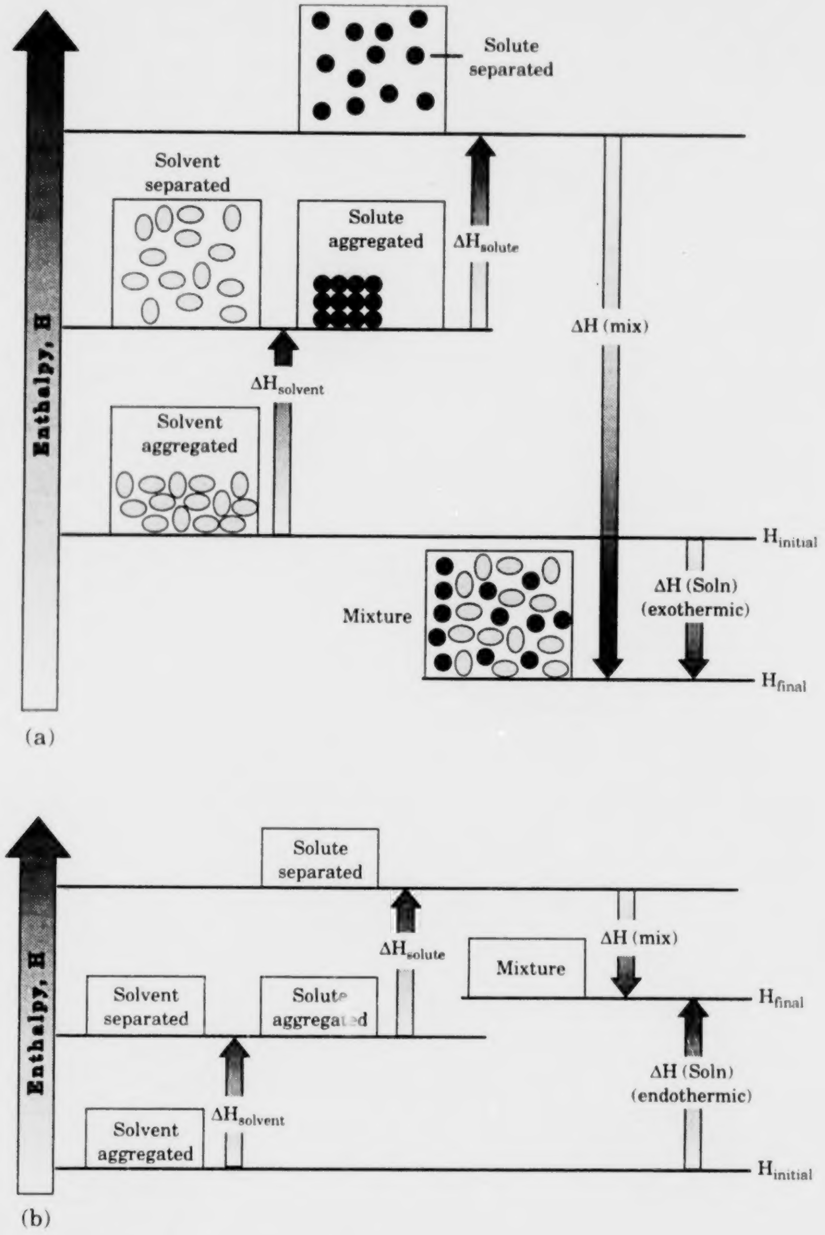

Enthalpy, H
Solute separated
Solvent separated
Solute aggregated
ΔH_{solute}
ΔH (mix)
$\Delta H_{solvent}$
Solvent aggregated
$H_{initial}$
Mixture
ΔH (Soln) (exothermic)
H_{final}
(a)
Enthalpy, H
Solute separated
ΔH_{solute}
ΔH (mix)
Mixture
H_{final}
Solvent separated
Solute aggregated
ΔH (Soln) (endothermic)
Solvent aggregated
$\Delta H_{solvent}$
$H_{initial}$
(b)

Figure 5. The conditions for dissolving.

Pattern Recognition

What Dissolves and What Does Not: Some General Rules. Charts similar to this appear in most chemistry texts.

What Dissolves and What Does Not: Some General Rules

After making a large number of observations on the solubility of compounds, chemists have devised some helpful "rules" to determine whether an ionic compound is soluble or insoluble in water. These generalizations are listed in the following chart.

The chart is easy to read. For example, it tells us that compounds containing ammonium ion (NH_4^+) are soluble (no exceptions). This means that NH_4Cl, $(NH_4)_2S$, $(NH_4)_3PO_4$, and so on are all soluble. It also tells us that compounds containing hydroxide (OH^-) are insoluble, except the alkali metals, calcium, barium, and strontium. For example, magnesium hydroxide, $Mg(OH)_2$, is insoluble, but calcium hydroxide, $Ca(OH)_2$, is soluble. With a little practice, students can make predictions about solubility of ionic compounds.

Compounds containing	Are	Except
Alkali metal ions (Li^+, Na^+, K^+, Rb^+, Cs^+)	Soluble	
Ammonium (NH_4^+)	Soluble	
Nitrate (NO_3^-)	Soluble	
Chlorate (ClO_3^-)	Soluble	
Perchlorate (ClO_4^-)	Soluble	
Acetate ($C_2H_3O_2^-$)	Soluble	
Chloride (Cl^-), Bromide (Br^-), Iodide (I^-)	Soluble	Lead (Pb^{2+}), Silver (Ag^+), and Mercury(I) (Hg_2^{2+}), NH_4^+
Sulfate (SO_4^{2-})	Soluble	Strontium (Sr^{2+}), Barium (Ba^{2+}), Mercury(I) (Hg_2^{2+}), Mercury(II) (Hg^{2+}), Lead (Pb^{2+}), NH_4^+
Hydroxide (OH^-)	Insoluble	Alkali metals, Ca^{2+}, Ba^{2+}, and Sr^{2+}
Sulfite (SO_3^{2-}), Chromate (CrO_4^{2-})	Insoluble	Alkali metals and NH_4^+
Carbonate (CO_3^{2-}), and Phosphate (PO_4^{3-})	Insoluble	Alkali metals and NH_4^+
Sulfide (S^{2-})	Insoluble	Alkali metals, Beryllium (Be^{2+}), Magnesium (Mg^{2+}), Calcium (Ca^{2+}), Strontium (Sr^{2+}), Barium (Ba^{2+}), and NH_4^+

Common Student Misconceptions

1. **"Like always dissolves like."**

   This is only a generalization and there are exceptions. For example, acetic acid is totally miscible in water due to the formation of hydrogen bonds, but it is also soluble in nonpolar solvents such as benzene and carbon tetrachloride.

2. **"The solubility of a solid in a liquid increases with an increase in temperature."**

   This is not always true. For example, sodium sulfite, calcium acetate, and lithium sulfate have solubilities that decrease with an increase in temperature.

3. **"Insoluble compounds do not dissolve."**

   This is not really true. All ionic compounds, even those that we classify as "insoluble" dissolve to some slight extent in water.

4. **"Hydration means the attachment of any solvent molecules to solute molecules."**

   Hydration is used only when the solvent is water. Solvation is the term used to describe any solvent molecule that surrounds and becomes attracted to solute molecules.

Problem Solving

Calculations

1. The solubility of $PbCrO_4$ is 1.34×10^{-7} mol/L at 25°C. Calculate the value of K_{sp}.

   Solution:

   $$PbCrO_4(s) \rightleftharpoons Pb^{2+}(aq) + CrO_4^{2-}(aq)$$
   $$K_{sp} = [Pb^{2+}][CrO_4^{2-}]$$
   $$K_{sp} = (1.34 \times 10^{-7})(1.34 \times 10^{-7})$$
   $$K_{sp} = 1.80 \times 10^{-14}$$

2. The K_{sp} for AgCl is 1.7×10^{-10} at 25 °C. What is the solubility of AgCl?

   Solution: Let's represent the concentration of Ag^+ as "x". Since there will be an equal concentration of Cl^-, this can be represented as "x" also. Then,

   $$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$$
   $$K_{sp} = [Ag^+][Cl^-]$$
   $$K_{sp} = [x][x]$$
   $$K_{sp} = x^2$$
   $$1.7 \times 10^{-10} = x^2$$
   $$x = \sqrt{1.7 \times 10^{-10}}$$
   $$x = 1.3 \times 10^{-5} \text{ mol/L}$$

Decision Making

1. Taking both entropy and enthalpy into consideration, explain why a gas is less soluble in water, whereas certain solids are more soluble in warm water. *[When a gas dissolves in water it forms weak bonds to the water and occupies less space. It also has slightly more order than it would as a free gas. Thus dissolving of a gas is exothermic (bond forming) and decreases entropy. Solids also generally dissolve exothermically but become considerably disordered as they move freely in the water. Since processes are favored by increased entropy and lower energy, solids tend to dissolve in water. Gases dissolving are favored by energy but not by entropy, so lower solubility is expected.]*

   $$\Delta G = \Delta H - T\Delta S$$
   $$\Delta G < 0 \text{ if } \Delta H < 0 \text{ and } \Delta S < 0$$

   (Note that increasing T makes ΔG smaller)

2. Refer to the accompanying table of solubility product constants to answer these questions:
   a. Although the ionic solids listed in the table are considered insoluble, which compound is the least soluble? *[CdS has the smallest Ksp and is therefore least soluble.]* Which is the most soluble? *[$CaSO_4$ is largest, thus most soluble.]*
   b. Which three compounds are the most insoluble? *[PbS, CdS, ZnS have the lowest K_{sp}'s.]* Does this suggest a generalization? *[Sulfides in general have low solubility.]*

NOTE: When a different number of ions per mole of slightly soluble salt form, the magnitude of the solubility product constant is not an immediately comparable value to determine relative solubility. For example, the K_{sp}'s of $CaCO_3$ and CaF_2 are 4.95×10^{-9} and 1.61×10^{-10}, respectively, but their molar solubilities are 7.04×10^{-5} and 3.43×10^{-4}, respectively.

Formula	K_{sp}	Formula	K_{sp}
AgBr	7.7×10^{-13}	$CaSO_4$	6.1×10^{-5}
AgCl	1.7×10^{-10}	$MgCO_3$	1×10^{-5}
AgCN	2×10^{-12}	MnS	1.4×10^{-15}
AgI	8.3×10^{-17}	$PbCO_3$	1.6×10^{-13}
$BaCO_3$	4.9×10^{-9}	$PbCrO_4$	1.8×10^{-14}
$BaSO_4$	1.5×10^{-9}	$PbSO_4$	1.9×10^{-6}
$CdCO_3$	2.5×10^{-14}	PbS	7×10^{-28}
CdS	1×10^{-28}	$SrSO_4$	2.8×10^{-7}
$CaCO_3$	4.8×10^{-9}	ZnS	4.5×10^{-24}

c. Certain toxic metal ions, including lead and mercury, are precipitated as a sulfide and then buried in an EPA approved landfill. Why is this an acceptable way to treat these toxic substances? *[Because these sulfides have such low solubility it is unlikely that they would dissolve and leach into the environment.]*

Graph Interpretation

Refer to Figure 6 to answer the following questions:

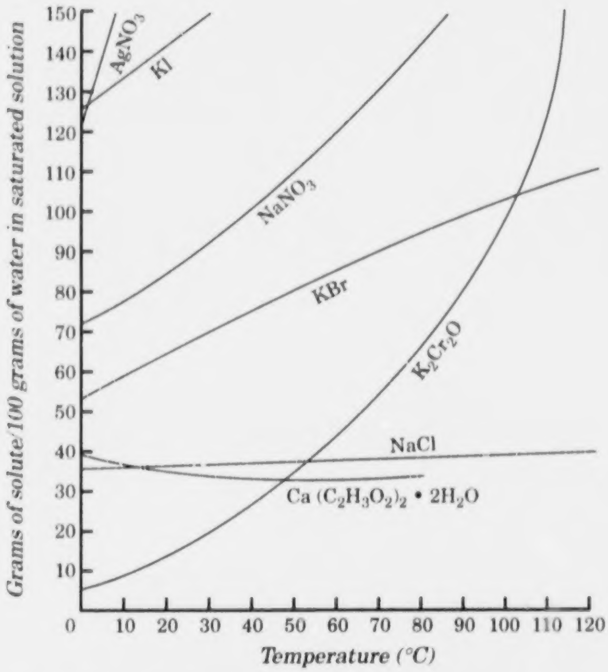


Figure 6. Solubilities of ionic compounds in water with respect to temperature.

1. Which substance decreases in solubility with increasing temperature? $[Ca(C_2H_3O_2) \cdot 2\ H_2O]$
2. Which substance has a solubility that is least affected by a change in temperature? [NaCl]

Additional Problem Solving

1. Suppose you were given a bottle of a solution that contained either magnesium, mercury(II), or barium ions. If you only had solutions of sodium hydroxide, sodium sulfate, and sodium chromate in your laboratory, devise a scheme to determine which ion is present in the mystery solution. You may find the chart "What Dissolves and What Does Not" (see Pattern Recognition) useful. *[Magnesium sulfate is soluble; barium hydroxide is soluble.]*
2. If you had a solution that contained a mixture of silver ions and strontium ions, devise four reactions that would allow you to precipitate one ion but leave the other in solution. *[Combine with hydroxide, sulfate, sulfide, halide ions.]*

HISTORY: ON THE HUMAN SIDE

Drinking Water

When considering the topic of solubility and solutions, one inevitably must discuss the role of water as the 'universal' solvent. While this property is necessary to sustain life, impurities can be dissolved or suspended in water and are often dangerous or even fatal to humans. The earliest records of water treatment and purification have been traced to approximately 2000 B.C. in the Sus'ruta Samhita, a body of Indian Sanskrit medical knowledge. Although not written down until 400 A.D., this oral tradition described a process to purify water by boiling over a fire, or being heated in the sun, or by dipping a heated iron into it; another way was to purify by filtration through sand and coarse gravel. During the period from the 15th to 13th centuries B.C. the Egyptians developed an apparatus for filtering and siphoning fluids and recorded images of them on the walls of the tombs of Amenophis II and Ramses II. By the eighth century A.D., Arabian alchemists had developed methods of distillation that provided a safer product than that of filtration.

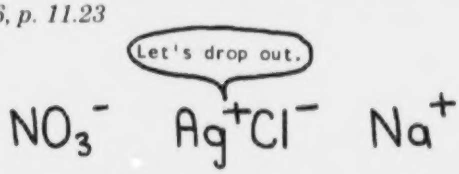
Most of Europe did not begin to recognize the need for water purification until the end of the Middle Ages. Sir Francis Bacon conducted numerous experiments on filtration, boiling, distillation, and clarification by coagulation. In 1685 an illustrated description of sand filters was published by Lucas Antonius Portius, an Italian physician. Portius had participated in efforts at mass sanitation for soldiers fighting the Austro-Turkish War of 1685. He proposed straining and sedimentation followed by multiple filtration through sand. Nevertheless, most Europeans continued to get their drinking water from streams, rivers, and rainwater collecting bins. They then brought these simple methods to the new world with them.

Until the end of the 19th century, Americans obtained drinking water from local ponds, wells, and rainwater holding tanks. Waste water was discarded in dry wells, leaching cesspools (pits lined with broken stones), or was simply dumped on the ground. Human defecation was usually collected in receptacles lined with brick or stone that were periodically emptied. However by the mid 1880s, flush toilets were beginning to find common use in urban settings. As a result, municipal sewer systems were constructed to carry the sewage. Until 1909 these sewer wastes were released without treatment into streams and lakes from which either the same urban center or others drew their drinking water. It was believed that rivers and large bodies of water would purify themselves.

As a result of disease epidemics that were striking urban centers at the turn of the century through the 1920s, urban centers began treating both the raw sewage from household wastes as well as chlorinating drinking water. Thus water is cleaned twice, once before being used, and again after use. The before-use cleaning, known as water treatment, takes place at municipal treatment plants. After its use, municipal water must again be treated, this time at a sewage treatment plant.

HUMOR: ON THE FUN SIDE

1. Message on a T-shirt: ***Front***: If you're not part of the solution...
   Back: You're part of the precipitate.
2. When students are deficient in note taking, ask them to precipitate some graphite.
   CHEM 13 NEWS, December 1976, p. 11.23
3. *CHEM 13 NEWS, January 1983, p. 11*



4. **THE GREEN LITTLE CHEMIST**

A green little chemist
On a green little day,
Mixed some green little chemicals
In a green little way.
The green little grasses
Now tenderly wave,
O'er the green little chemist's
Green little grave.

CHEM 13 NEWS, February 1981, p. 10

5. Word Search (see *Appendix* for master copy)

N N K N U F W J D Y I W J Q L V H
W P Z D C O T ~~W O L L E Y~~ G M A T
L Q ~~W H I T E~~ H K C R N R A D C E
P S O C X B K M S U W V E C S V T
~~K S P~~ V V E Y T T N U U D R C C A
M R T C L X X A R O A S E Z S X T
Y X V V F E R C I M L G E R I R I
X B S X O E O G Q M C Q ~~R A L O P~~
H K N G P B X B P O Z M H A V Y I
C U ~~A M M O N I U M~~ X T T M E J C
D J E L P J B Y Q P D Y W I R B E
~~Y T I L I B U L O S~~ X C L B L F R
X A S A C A O W O E S N G F E U P

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Ability of a substance to dissolve in a given quantity of solvent.
2. Solid formed by reaction in aqueous solution.
3. Color of solid formed when solutions of silver nitrate and sodium chloride are mixed.
4. Color of solid formed when solutions of lead nitrate and potassium iodide are mixed.
5. Ionic solids dissolve in this kind of solvent.
6. Solubility of most solids increases as this factor is increased.
7. Symbol for solubility product constant.
8. For PbI_2, Ksp = $[Pb^{2+}]^x [I^-]^y$. The sum of x + y is this number.
9. Compounds containing this polyatomic cation are soluble.
10. Halides of this metal are used in photography.

Answers: 1. SOLUBILITY 2. PRECIPITATE 3. WHITE 4. YELLOW 5. POLAR 6. TEMPERATURE 7. KSP 8. THREE 9. AMMONIUM 10. SILVER

6. See cartoons at end of module.

Links/Connections

WITHIN CHEMISTRY

1. Equilibrium
2. Analytical chemistry
3. Organic chemistry (especially organic solvents)
4. Chemical bonding

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

1. **Environmental Sciences**
    Air pollution
    Water pollution
    Acid rain
    Eutrophication in lakes and rivers (from dissolved nitrates and phosphates)
2. **Geochemistry**
    Formation of sink holes by solution of underground limestone
    Crystal formation
3. **Biochemistry**
    Biochemical reactions
    Fat-solubility of DDT, other pesticides
4. **Medicine**
    Artificial kidneys for hemodialysis
    Kidney and gallstone formation
    Preparation of intravenous fluids
5. **Meteorology**
    Rain, snow formation

TO THE CONTEMPORARY WORLD

1. **The Process of Life Depends upon Solutions**
    a. Oxygen and carbon dioxide are carried in solution in blood.
    b. Nutrients from foods we eat (simple sugars, amino acids, water soluble vitamins, *etc.*) are absorbed as particles dissolved in water.
    c. Fat-soluble vitamins are ingested as part of fatty substances and are stored in fat in the body.
    d. Urea and other excretory products are water soluble.
    e. Bones and teeth are composed primarily of insoluble phosphates, which contribute to their strength and durability.
2. **Altering Solubilities Often Results in Health Problems**
    a. Bacteria in the urinary tract can hydrolyze urea, producing ammonia and changing urine's pH. This lowers the solubility of magnesium ammonium phosphate ($MgNH_4PO_4$). This mineral (called struvite) precipitates as one type of kidney stone and must be removed surgically or broken up by ultrasound.

   b. High-protein intake increases the production of uric acid and urates. When their concentration exceeds their solubility, they often precipitate in various joints of the body. This condition is known as gout.

   c. Water-soluble vitamins are excreted in urine and must be replenished daily, or deficiency symptoms may result. Fat-soluble vitamins (A, D, E) are stored in body fat. One generally has enough of these vitamins stored to last several years.

3. **Formation of Caves**. Caves are formed because most rocks (especially limestone, gypsum, and dolomite) are soluble in water. Calcium carbonate, the primary component of limestone, dissolves in water rich in carbon dioxide to form calcium hydrogen carbonate. When the concentration of carbon dioxide decreases due to temperature change, pressure change, *etc.* as found in caves, solid calcium carbonate precipitates to form stalactites and stalagmites.

4. **Boiler Scale**. The same process described in (3) produces a deposit of calcium carbonate in water pipes and in boilers. When water saturated with carbon dioxide seeps through the ground it reacts with limestone to form soluble calcium hydrogen carbonate. When it is heated, carbon dioxide is driven off and insoluble calcium carbonate forms.

5. **Acid Rain**. Acid rain reacts with monuments, statues, buildings, *etc.* composed primarily of calcium carbonate to form the more soluble calcium sulfate.

6. **Photography**. Insoluble silver halides (AgI, AgBr, AgCl) are used to manufacture photographic film and paper.

7. **Extraction of Magnesium from Sea Water**. Hydroxide ion is added to sea-water to precipitate magnesium hydroxide, thus separating it from the more abundant sodium ions. The magnesium hydroxide is neutralized with hydrochloric acid. The resulting solid magnesium chloride is melted; electrolysis of the molten salt gives magnesium metal.

Media

1. "Solutions: Ionic and Molecular," Coronet laser videodisc (also on videocassette). Includes discussion questions.

2. "Doing Chemistry," laser videodisc set.

   Side C. DMEX C21 "Equilibrium Crystallization"
   DMEX C22 "Supersaturation"
   DEMO C25 "Aqueous and Nonaqueous Solvents"
   EXPT C28 "Solution Formation"
   EXPT C20 "Effect of Temperature on Solubility"

   Extensive hard copy descriptions and worked out experiments are provided with the disk. From American Chemical Society, 1155-16th St., N.W., Washington, D.C. 20036.

3. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).

   a. For the Apple II computer running on ProDOS: AR 303

   b. For the Apple II computer: AP 310, AP 502

   c. For IBM PCs and PC-compatibles: PC 2501

References

Module developed by Robert Davis, Pat Owens, and Lee Summerlin, the Alabama team.

Laboratory activities, demonstrations, and charts in this module were adapted from the following sources:

Borgford, C., and Summerlin, L. (1988). *Chemical activities*. Washington, DC: American Chemical Society.

A variety of activities, with teaching tips and safety discussions, useful demonstrations or laboratory experiments for grades 7-12.

Schaff, J., Niedfeldt, K., and Brawders, J. (1966). *Semi-micro experiments for the chemstudy program*. Boston, MA: D.C. Heath.

Summerlin, L. (1981). *Chemistry for the life sciences*. New York, NY: Random House.

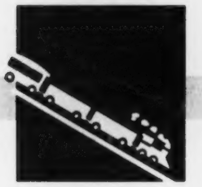
A text for the beginning college student planning a career in any health-related field, especially nursing and medical technology.

Summerlin, L., Borgford, C., and Ealy, J. (1988). *Chemical demonstrations, Volume II, Second Edition*. Washington, DC: American Chemical Society.

Summerlin, L., and Ealy, J. (1988). *Chemical demonstrations, Volume I, Second Edition*. Washington, DC: American Chemical Society.

Two good resources of demonstrations for teachers of chemistry and physics.

Appendix



- **Transparency Masters**
  1. Solution Set #1
  2. Suggested Chart for Reordering Solubility Observations
  3. Conditions for the Dissolving Process
  4. Word Search
- **Humor**

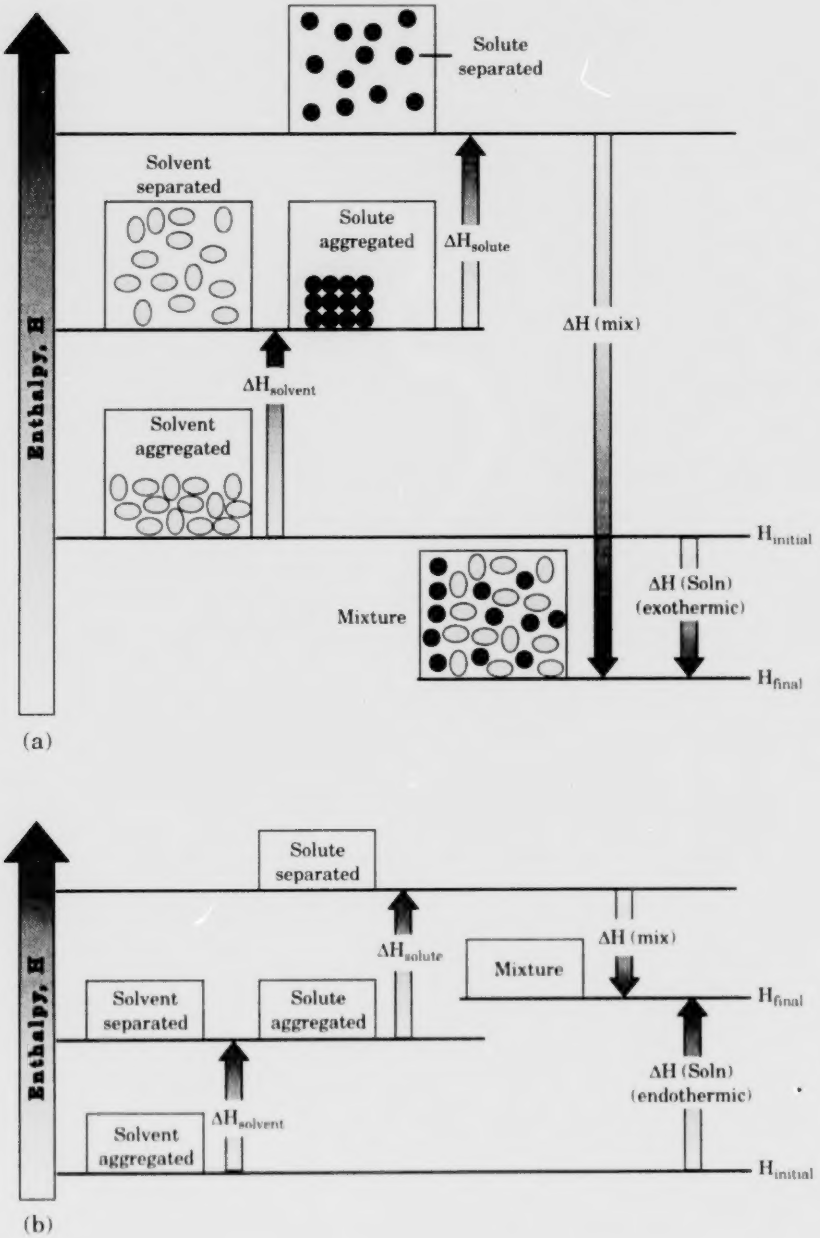
Solution Set #1

Set No. 1	K^+/Cl^- 6	Na^+/NO_3^- 5	Ag^+/NO_3^- 4	Na^+/I^- 3	Ba^{2+}/Cl^- 2
Ba^{2+}/NO_3^- 1					
Ba^{2+}/Cl^- 2					
Na^+/I^- 3					
Ag^+/NO_3^- 4					
Na^+/NO_3^- 5					

Suggested Chart for Reordering Solubility Observations

	Cl^-	CH_3COO^-	CO_3^{2-}	OH^-	NO_3^-	I^-	SO_4^{2-}
Ag^+							
Al^{3+}							
Ba^{2+}							
Co^{2+}							
Fe^{3+}							
K^+							
Mg^{2+}							
Na^+							
Sr^{2+}							

Conditions for the Dissolving Process


Enthalpy, H
Solute separated
Solvent separated
Solute aggregated
ΔH_{solute}
ΔH (mix)
$\Delta H_{solvent}$
Solvent aggregated
$H_{initial}$
Mixture
ΔH (Soln) (exothermic)
H_{final}
(a)
Enthalpy, H
Solute separated
ΔH_{solute}
ΔH (mix)
Mixture
H_{final}
Solvent separated
Solute aggregated
ΔH (Soln) (endothermic)
$\Delta H_{solvent}$
Solvent aggregated
$H_{initial}$
(b)

Word Search

N	N	K	N	U	F	W	J	D	Y	I	W	J	Q	L	V	H
W	P	Z	D	C	O	T	W	O	L	L	E	Y	G	M	A	T
L	Q	W	H	I	T	E	H	K	C	R	N	R	A	D	C	E
P	S	O	C	X	B	K	M	S	U	W	V	E	C	S	V	T
K	S	P	V	V	E	Y	T	T	N	U	U	D	R	C	C	A
M	R	T	C	L	X	X	A	R	O	A	S	E	Z	S	X	T
Y	X	V	V	F	E	R	C	I	M	L	G	E	R	I	R	I
X	B	S	X	O	E	O	G	Q	M	C	Q	R	A	L	O	P
H	K	N	G	P	B	X	B	P	O	Z	M	H	A	V	Y	I
C	U	A	M	M	O	N	I	U	M	X	T	T	M	E	J	C
D	J	E	L	P	J	B	Y	Q	P	D	Y	W	I	R	B	E
Y	T	I	L	I	B	U	L	O	S	X	C	L	B	L	F	R
X	A	S	A	C	A	O	W	O	E	S	N	G	F	E	U	P

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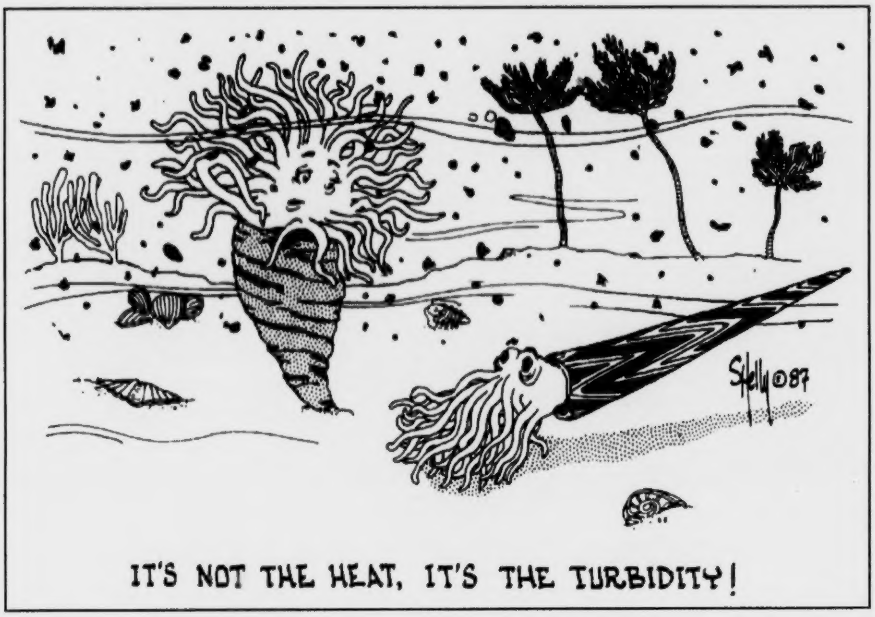
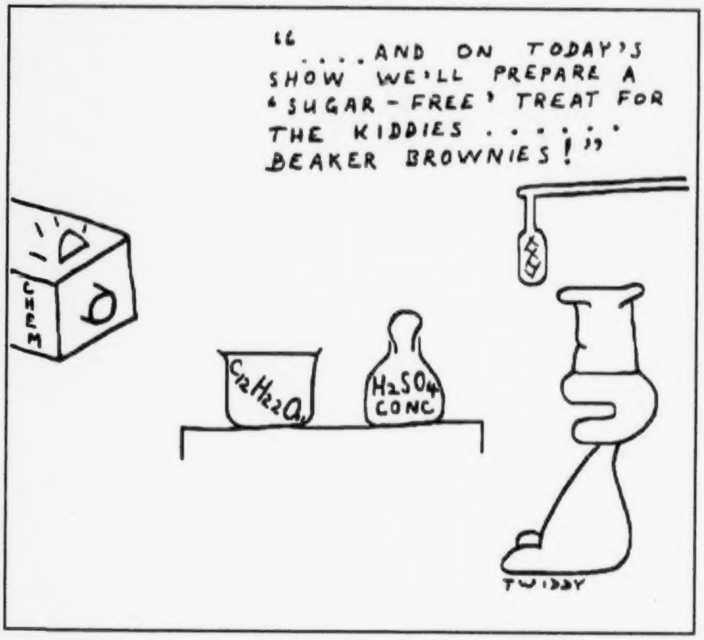

Shelly ©87
IT'S NOT THE HEAT, IT'S THE TURBIDITY!

Cartoon by Shelly Fischman. Reprinted with permission.

WHEN I TOLD YOU TO WASH YOUR PRECIPITATE, I MEANT...

Used by permission of Veritex Publishing Co.


"....AND ON TODAY'S SHOW WE'LL PREPARE A 'SUGAR-FREE' TREAT FOR THE KIDDIES ......BEAKER BROWNIES!"
CHEM
$C_{12}H_{22}O_{11}$
H_2SO_4 CONC
TWIDDY

CHEM 13 NEWS, February 1984, p. 16. Reprinted with permission.

A SourceBook Module

Version 2.0 1994

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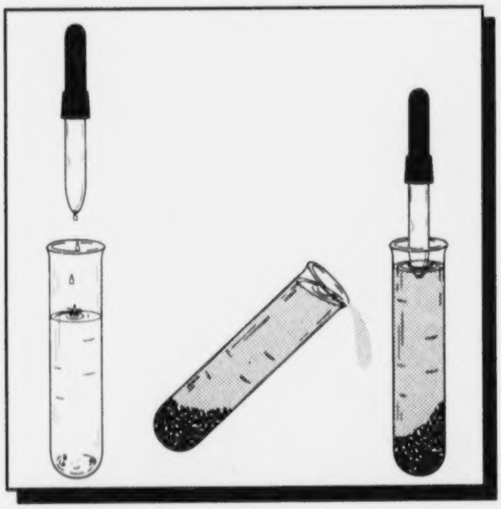
INORGANIC QUALITATIVE ANALYSIS



ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

From the crime lab to the water treatment plant, from environmental testing to agricultural soil testing to environmental quality determination, society asks chemistry to analyze materials and determine what and how much is contained in them.

Analysis is one of the central branches of chemistry. It is conveniently divided into qualitative analysis, which concentrates on what materials are present in a sample, and quantitative analysis, which is concerned with how much of a substance is present in a sample. In practice, qualitative and quantitative methods go hand-in-hand.

For many reasons, inorganic qualitative analysis laboratory has almost disappeared from the curriculum. In the first place, the modern practice of chemical analysis is being done with instrumentation, and more traditional "wet" chemistry methods might be considered old fashioned. Secondly, traditional qualitative analysis schemes of the metal ions were carried out using H_2S gas, which is noxious and toxic. Even with the substitution of thioacetamide as a source of sulfide, the odor of the qualitative analysis laboratory gave chemistry buildings their characteristic "stink." Finally, qualitative analysis schemes were so complicated that the scale of the laboratory required it to be taught over a period of one to two semesters. Since it became increasingly difficult to coordinate this laboratory with a concept-centered lecture course, it became necessary to teach qualitative analysis as a stand-alone course. As the undergraduate chemistry curriculum has become more loaded, it becomes less practical to devote so much time to this subject.

Nevertheless, qualitative analysis has many defenders who point out its value and argue for its place in the chemistry curriculum. First of all, qualitative analysis schemes are an excellent source of descriptive chemistry information. As beginning chemistry courses are becoming more principle oriented, many chemists are concerned that chemistry students don't have a basic repertoire of reaction chemistry that illustrates chemical concepts. Secondly, the development and use of qualitative analysis schemes require the kind of logical reasoning that is sometimes used as a justification for science courses. Finally, qualitative analysis schemes allow the development or review of such topics as solubility product equilibria, kinetics, acid/base chemistry, and coordination complexes. For these reasons, inorganic qualitative analysis has undergone a resurgence in recent years.

In order to answer the safety criticisms leveled at the use of qualitative analysis, alternative analysis schemes have been developed using benzoate, hydroxide, and phthalate anions. There has also been an effort to "scale-down" the amounts of noxious chemicals using semimicro, and in recent years, small-scale techniques. If the users of *SourceBook* want to carry out an extended qualitative analysis instruction using a traditional or more modern metal ion scheme, references can be found at the end of this module.

This module takes the view that qualitative analysis can play a useful role in the high school chemistry curriculum if it meets several criteria. (1) Qualitative analysis should be limited in scope so that it can be taught as a module or part of other modules. (2) Laboratory activities should be done with small amounts of chemicals and involve chemicals that can be handled safely. (3) The emphasis of a qualitative analysis scheme should be on the development of the scheme and its logic rather than exclusively on the use of a scheme developed by others.

PLACE IN THE CURRICULUM

As envisioned here, this module could be used as a review of chemical ideas developed during a high school chemistry course. As such it should be used at the end of a one-year course in chemistry, or during a second-year course. Many of the individual activities could also be used to supplement other modules. For example, some of the reactions in the qualitative analysis laboratory could be used to illustrate solubility and acid/base equilibrium. This module could also be used as an opportunity for students to develop and practice the logical processes used in chemistry.

CENTRAL CONCEPTS

1. Chemical substances can be identified (specified) from their chemical and physical characteristics and interactions.
2. A scheme or sequence of procedures can be used to identify the presence or absence of a chemical substance in combination with others.

RELATED CONCEPTS

1. Acids and bases are chemical species that exhibit distinctive sets of observable properties.
2. Electrons occupy specific energy levels in an atom. These energy levels are quantized and are responsible for emission spectra.
3. Reaction rates are influenced by factors such as the nature of the reactants, presence of a suitable catalyst, temperature, concentration of the reactants, and available surface area of solid or liquid reactants, or pressure for gases.
4. Many chemical reactions can be classified, at the simplest level, according to general kinds of change involved:

$A + B \rightarrow AB$	Synthesis
$AB \rightarrow A + B$	Decomposition
$AB + C \rightarrow AC + B$	Replacement
$AB + CD \rightarrow AD + CB$	Double replacement

   Many chemical reactions involve the transfer of electrons between reacting chemical species.

5. When a solute is only slightly soluble, the extent of solubility can be quantitatively expressed as the solubility product constant.
6. The solubility of a substance can be altered by common ion effect, pH, temperature change, formation of a complex, *etc.*
7. Accumulated data allow one to predict the degree of solubility of many substances, *i.e.*, solubility rules.

RELATED SKILLS

1. Name simple ionic compounds.
2. Use flame tests to identify active metals.
3. Use indicators to determine the pH of solutions.
4. Write balanced chemical equations for chemical reactions.
5. Hypothesize chemical reactions from reaction types.
6. Use a centrifuge to separate a precipitate from a solution.
7. Use basic separation techniques.
8. Utilize solubility rules.
9. Utilize basic chemical tests to identify common gases and other chemicals.

PERFORMANCE OBJECTIVES

After completing their study of qualitative analysis, students should be able to:

1. Analyze a sample for the presence of a specific ion, given testing procedures.
2. Develop an analysis scheme for a set of ionic substances.
3. Use comparative reaction information to identify a precipitate.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Qualitative Analysis Laboratory I: Developing a Qualitative Analysis Scheme for the Cations Zn^{2+}, Ca^{2+}, Cu^{2+}, Al^{3+}, Co^{2+}

Introduction

Solutions often contain unknown ions. Fortunately, patterns of chemical reactivity and solubility allow for the isolation and identification of ions in a solution. This identification activity is called qualitative analysis. In this activity you will devise your own identification scheme for five cations.

Purpose

To study some of the properties and patterns of reactivity that allow for the separation and identification of certain ions in a solution.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Be careful to avoid contaminating the test solutions and stock solutions.

Procedure

Obtain approximately 10 mL of 0.1 M solutions Zn^{2+}, Ca^{2+}, Cu^{2+}, Al^{3+} and Co^{2+} in separate clean, dropper bottles. (Each of these solutions was prepared from the nitrate salt of the cation.) You are to subject each of these solutions to the tests outlined below and record your observations in the table provided.

Tests

1. Put 10 drops test solution (Zn^{2+}) into a 75-mm test-tube. Using pH paper, test the pH. If needed, add 3 M NH_3 or 3 M HNO_3 dropwise until the pH is neutral, *i.e.*, between 6 and 8. Now add 7 drops 0.3 M Na_3PO_4.
2. To 10 drops fresh test solution (Zn^{2+}) add 3 M HNO_3 until the solution is strongly acidic, *i.e.*, pH<3. Add 7 drops 0.2 M $K_3Fe(CN)_6$, potassium ferricyanide.
3. To 10 drops fresh test solution (Zn^{2+}) add 6 M NaOH until the solution is just basic, *i.e.*, pH >8. Add 10 drops 6 M NaOH. If a persistent precipitate forms, mix the sample and centrifuge. Decant the solution and save any precipitate formed.
4. Wash the precipitate twice with about 5 mL of distilled water. Add 10 drops 6 M NH_3 and mix thoroughly, then centrifuge for 1 min. Observe.
5. To 10 drops of fresh test solution (Zn^{2+}) add 6 M NaOH until the solution is just basic. Add 10 drops NaOH. Mix, then centrifuge if a precipitate develops. Decant. If there is a precipitate, wash it twice with distilled water and centrifuge. Add 15 drops 3 M HNO_3 to the precipitate and mix.

   Write your observations (precipitates, colors) in the table that follows.
6. Repeat Steps 1-5 for the remaining 4 cations.
7. Thoroughly wash your hands before leaving the laboratory.

Reagent	Cation				
	Zn^{2+}	Ca^{2+}	Cu^{2+}	Al^{3+}	Co^{2+}
Test 1 PO_4^{3-}					
Test 2 $Fe(CN)_6^{3-}$					
Test 3 OH^-					
Test 4 NH_3					
Test 5 H^+					

Data Analysis and Concept Development

From your data complete the equations listed below. You will be speculating, but your speculation will be accurate in 90% of the cases if you use the following guidelines. (1) Assume that most reactions are a simple combination of cations and anions. (2) If the product is dissolved, it is ionic. (3) If the product is solid, it is neutral. If you observed no change when the particular reagents were mixed, write N.R. for "no reaction."

1. $Zn^{2+} + PO_4^{3-} \rightarrow$ ________
2. $Ca^{2+} + PO_4^{3-} \rightarrow$ ________
3. $Cu^{2+} + PO_4^{3-} \rightarrow$ ________
4. $Al^{3+} + PO_4^{3-} \rightarrow$ ________
5. $Co^{2+} + PO_4^{3-} \rightarrow$ ________
6. $Zn^{2+} + Fe(CN)_6^{3-} \rightarrow$ ________
7. $Ca^{2+} + Fe(CN)_6^{3-} \rightarrow$ ________
8. $Cu^{2+} + Fe(CN)_6^{3-} \rightarrow$ ________
9. $Al^{3+} + Fe(CN)_6^{3-} \rightarrow$ ________
10. $Co^{2+} + Fe(CN)_6^{3-} \rightarrow$ ________
11. $Zn^{2+} + OH^- \rightarrow$ ________ $\xrightarrow{NH_3}$ ________
    $Zn^{2+} + OH^- \rightarrow$ ________ $\xrightarrow{H^+}$ ________
12. $Ca^{2+} + OH^- \rightarrow$ ________ $\xrightarrow{NH_3}$ ________
    $Ca^{2+} + OH^- \rightarrow$ ________ $\xrightarrow{H^+}$ ________

13. $Cu^{2+} + OH^- \rightarrow$ ________ $\xrightarrow{NH_3}$ ________

$Cu^{2+} + OH^- \rightarrow$ ________ $\xrightarrow{H^+}$ ________

14. $Al^{3+} + OH^- \rightarrow$ ________ $\xrightarrow{NH_3}$ ________

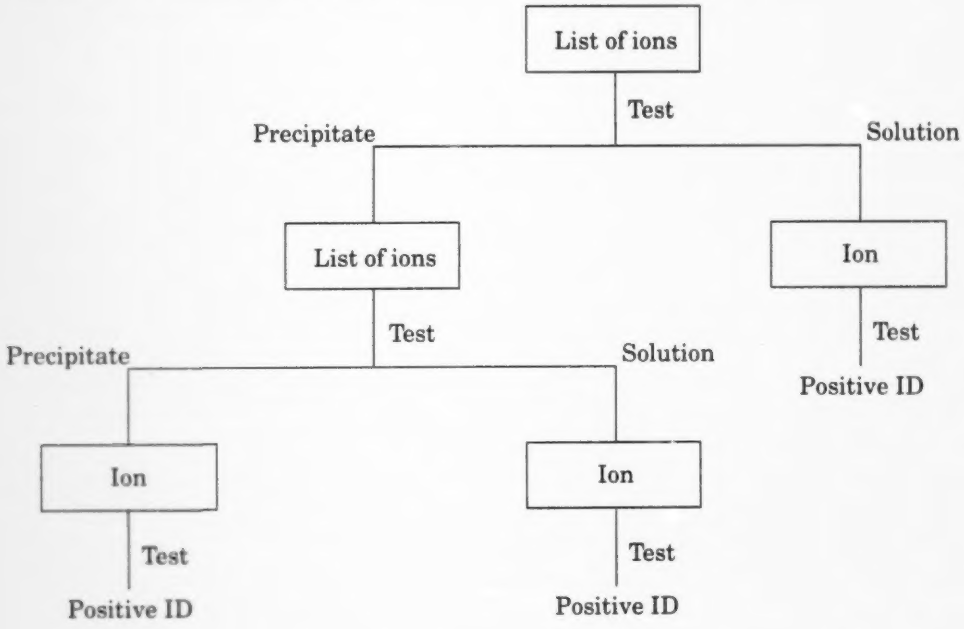
$Al^{3+} + OH^- \rightarrow$ ________ $\xrightarrow{H^+}$ ________

15. $Co^{2+} + OH^- \rightarrow$ ________ $\xrightarrow{NH_3}$ ________

$Co^{2+} + OH^- \rightarrow$ ________ $\xrightarrow{H^+}$ ________

Implications and Applications

1. Assume you were given a solution that was said to contain Ca^{2+} cations only. Describe how you would confirm or deny the presence of this cation in the solution. Explain your reasoning.
2. Assume you were given a solution that was said to contain Zn^{2+} and/or Al^{3+} cations only. Describe how you would determine if neither, only one or both of these ions were present in solution. Explain your reasoning.
3. Assume you were given a solution that was said to contain Cu^{2+} and/or Co^{2+} cations. Describe how you would determine if neither, only one or both of these ions were present in solution.
4. Assume you were given a solution that was said to contain Zn^{2+}, Ca^{2+}, Cu^{2+}, Al^{3+}, and/or Co^{2+} cations only. Describe how you would confirm or deny the presence of each of these ions in the solution. Explain your reasoning. (Hint: You are encouraged to design an analysis scheme—a flow chart—such as the one shown in the following example.)


List of ions
Test
Precipitate
Solution
List of ions
Ion
Test
Precipitate
Solution
Test
Positive ID
Ion
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Positive ID

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Qualitative Analysis Laboratory I: Developing a Qualitative Analysis Scheme for the Cations Zn^{2+}, Ca^{2+}, Cu^{2+}, Al^{3+}, Co^{2+}

NOTE: A more guided alternative to the suggested laboratory can be found in Dreyfus Microscale Lab: Qualitative Analysis #36 (Chen & Ingels, 1987).

Major Chemical Concept

Cations in solution can be separated and identified by characteristic chemical changes. These changes can be observed by precipitate formation, color change, or formation of gases.

Level

General and Honors

Expected Student Background

1. Experience with observational evidence for chemical change.
2. Experience in chemical formula writing and knowledge of common ion charges.
3. Ability to balance chemical equations.
4. Some knowledge of solubility patterns for selected compounds. (A list of "solubility rules" can be found in most basic chemistry books.)

Time

The first set of tests will require two or three 50-min laboratory periods. The *Data Analysis and Concept Development* section can be assigned as homework. *Implications and Applications* will require one or two periods.

Safety

Read the *Safety Considerations* in the *Student Version*.

1. Students should be cautioned that the solutions are toxic. Some are caustic and should be handled with care.
2. Students should wash their hands prior to leaving the laboratory.
3. They should also wash their hands if their skin feels "slippery." This sensation indicates they have come in contact with NaOH.
4. Safety goggles should be worn while performing these activities.

Materials (For 24 students working in pairs)

Nonconsumables

12	Test-tube racks
120	Test-tubes, 75-mm
72	Medicine droppers
60	Stirring rods
2	Vials universal pH paper
4	Centrifuges
5	Dropper bottles per group

Consumables

3 M Ammonia, NH_3, 150 mL [30 mL 15 M (conc.) NH_3 per 150 mL solution]
3 M Nitric acid, HNO_3, 250 mL [47 mL 16 M (conc.) HNO_3 per 250 mL solution]
0.2 M Potassium ferricyanide, $K_3Fe(CN)_6$, 150 mL (9.9 g $K_3Fe(CN)_6$ per 150 mL solution) *NOTE:* $K_3Fe(CN)_6$ is called potassium ferricyanide—*not* ferrocyanide. The correct name for this compound is potassium hexacyanoferrate(III).
0.3 M Trisodium phosphate, Na_3PO_4, 150 mL (17.1 g $Na_3PO_4 \cdot 12H_2O$) per 150 mL solution)
6 M Sodium hydroxide, NaOH, 150 mL (36 g NaOH per 150 mL solution)
6 M Ammonia, NH_3, 150 mL (60 mL 15 M (conc.) concentrated NH_3 per 150 mL solution)
0.1 M Zinc nitrate, $Zn(NO_3)_2$, 150 mL (4.5 g $Zn(NO_3)_2 \cdot 6H_2O$ per 150 mL solution)
0.1 M Calcium nitrate, $Ca(NO_3)_2$, 150 mL (3.5 g $Ca(NO_3)_2 \cdot 4H_2O$ per 150 mL solution)
0.1 M Copper(II) nitrate, $Cu(NO_3)_2$, 150 mL (3.6 g $Cu(NO_3)_2 \cdot 3H_2O$ per 150 mL solution)
0.1 M Aluminum nitrate, $Al(NO_3)_3$ 150 mL (5.6 g $Al(NO_3)_3 \cdot 9H_2O$ per 150 mL 0.01 M HNO_3 solution) [0.01 M HNO_3 (0.5 mL 3 M HNO_3 per 150 mL solution).]
0.1 M Cobalt(II) nitrate, $Co(NO_3)_2$, 150 mL (4.4 g $Co(NO_3)_2 \cdot 6H_2O$ per 150 mL solution)
pH Paper Range 0-13

Advance Preparation

1. Sets of solutions of each reagent should be prepared and placed in labeled dropper bottles in convenient locations in the laboratory. Students should be advised to return the dropper to its bottle immediately after using it. This will avoid contamination of the reagents.
2. The reading on semi-micro laboratory techniques (included in the *Appendix*) should be reproduced and given to students to read prior to the first laboratory period. The procedure assumes that students understand these techniques.

Pre-Laboratory Discussion

1. Review and answer student questions on the semi-micro techniques reading. Students should use the reading for reference as needed, while they are performing the procedures.
2. Review the safety precautions to be used in the laboratory.
3. Stress the importance of and procedures for keeping reagent samples and stock bottles uncontaminated.
4. Instruct students on safe disposal of wastes from the activities. It is suggested that the liquid wastes be collected in a large, centrally located bottle.

   (Consult section on *Waste Disposal* in *Appendix.*)

Teacher-Student Interaction

1. Students will start slowly and increase in speed with experience. Observe their procedures and correct them if the techniques are sloppy or leading to contamination of stock reagents or samples.

2. Examine student observation sheets to insure that their observations are consistent with normal expectations. If they have aberrant observations suggest that they repeat the test. Do not tell them the correct answers.
3. You might want to divide this laboratory into parts. Have students do some of the tests and then hold a discussion on their observations. Pool their data and discuss writing equations. You might also use these discussions to explore the chemistry of these materials. Use the *Notes on Tests and Equations* in the *Answers to Data Analysis* section as a basis for discussion of student observations. Be prepared to rerun a test as a demonstration.

Anticipated Student Results

Reagent	Cation				
	Zn^{2+}	Ca^{2+}	Cu^{2+}	Al^{3+}	Co^{2+}
Test 1 PO_4^{3-}	White ppt	White ppt	Blue ppt	White ppt	Purple ppt
Test 2 $Fe(CN)_6^{3-}$	Yellow ppt		Yellow-green ppt		Red ppt
Test 3 OH^-	Initial ppt disappears	White ppt, but disappears by 2nd wash	Blue ppt	Initial ppt disappears	Blue ppt fading to pink
Test 4 NH_3			ppt dissolves Dark blue solution		ppt remains
Test 5 H^+			ppt dissolves Pale green solution		ppt dissolves Pink solution

Answers to Data Analysis and Concept Development

Equations

1. $3Zn^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Zn_3(PO_4)_2(s)$
2. $3Ca^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$
3. $3Cu^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Cu_3(PO_4)_2(s)$
4. $Al^{3+}(aq) + PO_4^{3-}(aq) \rightarrow AlPO_4(s)$
5. $3Co^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Co_3(PO_4)_2(s)$
6. $3Zn^{2+}(aq) + 2Fe(CN)_6^{3-}(aq) \rightarrow Zn_3[Fe(CN)_6]_2(s)$
7. $Ca^{2+}(aq) + Fe(CN)_6^{3-}(aq) \rightarrow$ N.R.
8. $3Cu^{2+}(aq) + 2Fe(CN)_6^{3-}(aq) \rightarrow Cu_3[Fe(CN)_6]_2(s)$
9. $Al^{3+}(aq) + Fe(CN)_6^{3-}(aq) \rightarrow$ N.R.
10. $3Co^{2+}(aq) + 2Fe(CN)_6^{3-}(aq) \rightarrow Co_3[Fe(CN)_6]_2(s)$

11. $Zn^{2+}(aq) + OH^{-}(aq) \rightarrow \text{N.R.} \xrightarrow{NH_3} \text{N.R.}$
    $Zn^{2+}(aq) + OH^{-}(aq) \rightarrow \text{N.R.} \xrightarrow{H^+} \text{N.R.}$
12. $Ca^{2+}(aq) + 2OH^{-}(aq) \rightarrow \text{N.R. } \textit{or } Ca(OH)_2(s) \xrightarrow{H_2O} Ca^{2+}(aq) + 2OH^{-}(aq)$
    $Ca^{2+}(aq) + 2OH^{-}(aq) \rightarrow \text{N.R. } \textit{or } Ca(OH)_2(s) \xrightarrow{H^+} Ca^{2+}(aq)$
13. $Cu^{2+}(aq) + 2OH^{-}(aq) \rightarrow Cu(OH)_2(s) \xrightarrow{NH_3} Cu(NH_3)_4^{2+}(aq) + 2OH^{-}(aq)$
    $Cu^{2+}(aq) + 2OH^{-}(aq) \rightarrow Cu(OH)_2(s) \xrightarrow{H^+} Cu^{2+}(aq) + 2H_2O(l)$
14. $Al^{3+}(aq) + OH^{-}(aq) \rightarrow \text{N.R.} \xrightarrow{NH_3} \text{N.R.}$
    $Al^{3+}(aq) + OH^{-}(aq) \rightarrow \text{N.R.} \xrightarrow{H^+} \text{N.R.}$
15. $Co^{2+}(aq) + 2OH^{-}(aq) \rightarrow Co(OH)_2(s) \xrightarrow{NH_3} \text{N.R.}$
    $Co^{2+}(aq) + 2OH^{-}(aq) \rightarrow Co(OH)_2(s) \xrightarrow{H^+} Co^{2+}(aq) + 2H_2O$

Notes on the Tests and Equations

1. Students may point out that most solubility rules would indicate that Zn^{2+} and Al^{3+} form insoluble hydroxides. Some may have noticed precipitates appearing when OH^- is added. This precipitate disappears when additional OH^- is added. Both $Zn(OH)_2$ and $Al(OH)_3$ are amphoteric (*i.e.*, they are soluble in both strong base and strong acid). They form soluble ions when an excess of OH^- is present. For example:

   $$Zn^{2+}(aq) + 2OH^{-}(aq) \rightarrow Zn(OH)_2(s)$$
   $$Zn(OH)_2(s) + 2OH^{-}(aq) \rightarrow Zn(OH)_4^{2-}(aq)$$

   In a similar manner $Al(OH)_4^-$ is formed. Students will probably not be able to write an equation for the complex ion formation and will merely indicate no reaction since that is consistent with their observations. The point of the test is that it allows the separation of Zn^{2+} and Al^{3+} from the other ions. Whether or not you discuss complex formation by these ions depends upon the background of students.

2. The $Ca(OH)_2$ precipitate dissolves in the washings because although it is "insoluble," it is relatively more soluble than most "insoluble" substances. ($K_{sp} = 5.5 \times 10^{-6}$)

3. The $Cu(OH)_2$ precipitate is difficult to see until the solution is centrifuged. The green solution formed when the $Cu(OH)_2$ is acidified is very pale. Most students will probably observe no color. The dark blue solution formed when the ammonia is added to the $Cu(OH)_2$ is due to the formation of the complex ion $Cu(NH_3)_4^{2+}$. Students are unlikely to write the correct formula for the tetraammine copper complex, and might just write $CuNH_3^{\;2+}$. It is probably not important that they write the correct formula, although it does not do any harm just to tell them the correct formula.

   The equations for the competing equilibria follow:

   $$Cu(OH)_2(s) \rightleftharpoons Cu^{2+}(aq) + 2OH^{-}(aq) \qquad K_{sp} = 1.6 \times 10^{-19}$$
   $$Cu^{2+}(aq) + 4NH_3(aq) \rightleftharpoons Cu(NH_3)_4^{\;2+}(aq) \qquad K_f = 2 \times 10^{14}$$

Because the equilibrium constants are close to offsetting each other, the combined reaction:

$$Cu(OH)_2\,(s) + 4NH_3(aq) \rightleftharpoons Cu(NH_3)_4^{2+}(aq) + 2OH^-(aq)$$

will dissolve the insoluble $Cu(OH)_2$ with a small excess of NH_3. Similarly acid will dissolve the precipitate:

$$Cu(OH)_2\,(s) + 2H^+(aq) \rightleftharpoons Cu^{2+}(aq) + 2H_2O(l)$$

4. $Co(OH)_2$ can exist as either a blue or a pink precipitate. The blue form is believed to be more finely dispersed. It changes to the coarser pink form that is more stable. Air will oxidize $Co(OH)_2$ to $Co(OH)_3$, which is a black/brown precipitate. If students do their cobalt ion tests during the same period, using fresh precipitates, the presence of $Co(OH)_3$ can be ignored. As with copper ions, it is possible to form the amine complex of cobalt ions. The equations follow:

$$Co(OH)_2\,(s) \rightleftharpoons Co^{2+}(aq) + 2OH^-(aq) \qquad K_{sp} = 12.5 \times 10^{-16}$$

$$Co^{2+}(aq) + 4NH_3(aq) \rightleftharpoons Co(NH_3)_4^{2+}(aq) \qquad K_f = 8 \times 10^4$$

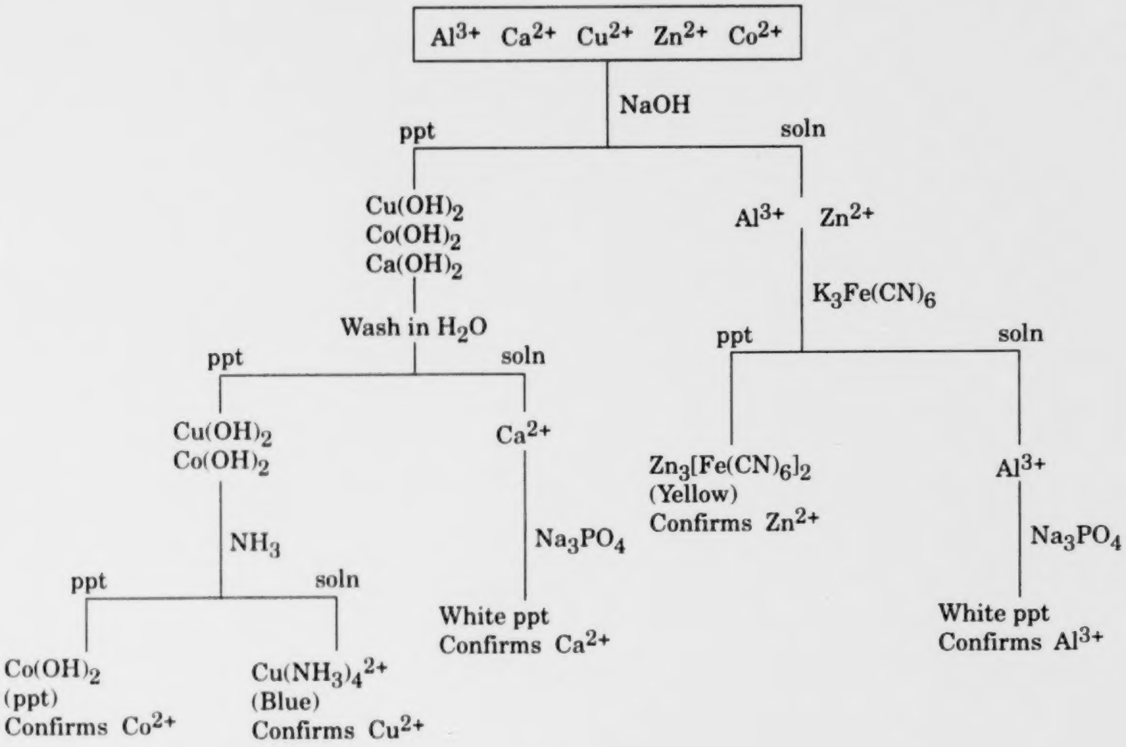
Because the relative sizes of the equilibrium constants favor the formation of insoluble $Co(OH)_2$ over $Co(NH_3)_4^{2+}$, the precipitate will not dissolve unless large excesses of ammonia are present. Thus, in contrast with $Cu(OH)_2$, $Co(OH)_2$ does not dissolve. This gives a basis for separating and identifying Cu^{2+} and Co^{2+}.

5. Adjusting the pH of solution is not much of a problem in doing these tests. However, if an unknown or a combination of ions used to try out separations is used, pH control becomes more critical. If the 0.3 M solutions are too concentrated to make the pH adjustments, suggest that students make dilute solutions for this purpose.

Answers to Implications and Applications

1. Add NaOH to the solution, and, if Ca^{2+} is present, a white precipitate will form. Wash the precipitate, save the solution liquid, and add Na_3PO_4 to the solution liquid. The formation of another white precipitate confirms the presence of Ca^{2+}.
2. Add potassium ferricyanide and nitric acid solutions to the solution. A yellow precipitate confirms the presence of Zn^{2+}. Centrifuge and decant the solution liquid into another test-tube, and add Na_3PO_4 solution. If a white precipitate forms, Al^{3+} is confirmed. If both of these tests produce the expected precipitate, then both ions are present.
3. Addition of NaOH to this solution will result in the formation of precipitates of both $Co(OH)_2$ and $Cu(OH)_2$. $Cu(OH)_2$ is soluble in concentrated NH_3, and the $Cu(NH_3)_4^{2+}$ ion forms a dark blue solution. If all the precipitate dissolves, then Co^{2+} was not present. If the $Cu(OH)_2$ dissolves and there is a dark precipitate remaining, then the presence of Co^{2+} is confirmed by that $Co(OH)_2$ precipitate.
4. See the following flow chart.

   Alternatives: Can dissolve $Co(OH)_2$ in acid and confirm with $K_3Fe(CN)_6$ or Na_3PO_4. Can use a portion of the unwashed NaOH precipitate to look for Ca^{2+} by dissolving in acid, precipitating Cu^{2+} and Co^{2+} with $K_3Fe(CN)_6$ and then using Na_3PO_4 to look for Ca^{2+}.


Al^{3+} Ca^{2+} Cu^{2+} Zn^{2+} Co^{2+}
NaOH
ppt
soln
$Cu(OH)_2$
$Co(OH)_2$
$Ca(OH)_2$
Wash in H_2O
ppt
soln
$Cu(OH)_2$
$Co(OH)_2$
NH_3
ppt
soln
$Co(OH)_2$
(ppt)
Confirms Co^{2+}
$Cu(NH_3)_4^{2+}$
(Blue)
Confirms Cu^{2+}
Ca^{2+}
Na_3PO_4
White ppt
Confirms Ca^{2+}
Al^{3+} Zn^{2+}
$K_3Fe(CN)_6$
ppt
soln
$Zn_3[Fe(CN)_6]_2$
(Yellow)
Confirms Zn^{2+}
Al^{3+}
Na_3PO_4
White ppt
Confirms Al^{3+}

Post-Laboratory Discussion

Discuss the reactions involved in the various tests with the class. Review the answers to the questions by having students submit their answers to the first three. Other ions may be the subject of similar questions raised to insure understanding of the chemical reactions and the laboratory techniques.

Extensions

While you may wish to include the analysis of an unknown as a learning assessment activity, such an activity could be used to support the knowledge gained from the laboratory work. The activity is included here.

Analysis of an Unknown

You will be given 10 mL of an unknown solution that might contain any or all of the cations: Zn^{2+}, Ca^{2+}, Cu^{2+}, Al^{3+}, and Co^{2+}. You are to analyze your solution in whatever manner you feel is appropriate to determine unequivocally which of the cations are present.

The laboratory report for this section should include the following:

1. **An experimental section**. This should be a step-by-step accounting of the activities conducted and the results observed. Include balanced ionic equations where appropriate.
2. **A conclusion section**. This should be a paragraph telling which cations you think were in your solution and telling why you reached this conclusion.

Notes to the Teacher

1. Use 0.15 M Al^{3+} in unknowns. The Al^{3+} test in our suggested Flow Chart is weak indicating that some Al^{3+} previously precipitated with NaOH. Make sure the unknown solutions are 0.1 M in each of the other metal ions used. Don't merely mix the test solutions since the concentrations will be too weak.
2. pH control called for in the tests is essential. Students should be reminded of this especially when doing the unknown.

Prediction and Naming of Precipitates

Following are two sets of solutions. Each set contains six solutions. One of these sets will be assigned to you. Mix small amounts of each sample with every other sample and observe any evidence of a chemical reaction. How many possible combinations of two solutions are possible? Write balanced equations for any chemical reactions you observe. Identify any precipitate formed.

Set 1		Set 2	
Na_2SO_4	$BaCl_2$	$FeCl_3$	NaOH
$Al_2(SO_4)_3$	$Ba(NO_3)_2$	$Co(NO_3)_2$	KOH
$Sr(NO_3)_2$	$AlCl_3$	$CoCl_2$	$NaNO_3$

Notes to the Teacher

Another laboratory activity, which can enhance the learning regarding prediction and naming of precipitates, involves combining of solutions of known compounds. Two sets of chemicals are offered as suggestions. Many other sets are possible. 0.1 M solutions of each substance should be provided in dropper bottles. Students might make a 6 x 6 grid for recording information about reaction results. Set 1 is used as an example. [A grid similar to the one below (without the expected results) can be photocopied and overlaid with transparent food wrap such as Saran™ wrap. A drop of each reagent in the vertical column can be placed on the appropriate squares in the rows, and then tested with a drop of appropriate reagent from the set along the top of the chart. This procedure is quick, easy to accomplish, and easy to interpret.]

	Na_2SO_4	$Al_2(SO_4)_3$	$Sr(NO_3)_2$	$BaCl_2$	$Ba(NO_3)_2$	$AlCl_3$
Na_2SO_4	——		White ppt	White ppt	White ppt	
$Al_2(SO_4)_3$		——	White ppt	White ppt	White ppt	
$Sr(NO_3)_2$	White ppt	White ppt	——			
$BaCl_2$	White ppt	White ppt		——		
$Ba(NO_3)_2$	White ppt	White ppt			——	
$AlCl_3$						——

Assessing Laboratory Learning

A laboratory practical to identify ions in an unknown solution is an effective way to assess learning (see *Extensions*).

Questions

1. Based on your laboratory experience, write balanced equations for each of the following pairs of reactants. Assume that chemical changes occur with each set of reactants.

$Al(NO_3)_3 + NaCl \rightarrow$

$CoCl_2 + NaOH \rightarrow$

$BaCl_2 + H_2SO_4 \rightarrow$

$Fe(NO_3)_3 + NaOH \rightarrow$

2. In which of the reactions in Question 1 would you be able to observe the formation of a precipitate?

A more challenging activity consists of providing students with a set of solutions labeled A, B, C, *etc*. Students are given the formulas of the chemicals in the solution and told to match the formula to the solution. Students make observations of reactions that occur when the solutions are mixed together. Many versions of this exercise exist—*e.g.*, Dreyfus Microscale Lab: Six Solutions #35 (Chen & Engels, 1987), but teachers can make up their own. Provide students with a solubility rules table and a table of common chemical tests (see *Appendix*).

DEMONSTRATIONS

CAUTION: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Flame Tests

Description

Cations can be identified by characteristic colors, when a compound containing the cation is heated in a burner flame. The purpose of this demonstration is to show the colors and to associate them with the cations in the compounds.

Materials

Use samples of $BaCl_2$, $CaCl_2$, $CuCl_2$, $SrCl_2$, NaCl, KCl.

Procedure

Details of setting up and carrying out of a flame test demonstration can be found in the *Alkali Metals* module.

Remarks

It would be appropriate to review the demonstration by testing "unknown" solutions and asking students to identify the cation present because of the color. Explain that these flame tests provide a method for qualitative analysis for cations, if known samples containing that cation can be obtained and tested.

For reference, expected colors of the cations used in this demonstration follow.

Ba^{2+}	Greenish-yellow	
Ca^{2+}	Brick-red	
Cu^{2+}	Green	
Sr^{2+}	Crimson	
Na^+	Yellow	
K^+	Violet	(This color can sometimes be enhanced by observing the flame through cobalt blue glass, which absorbs all but blue and violet light.)

Demonstration 2: Spectrum Demonstration

Purpose

To show how colored substances absorb characteristic wavelengths of light in the visible spectrum.

Materials

Sheet of diffraction grating, 6.5- x 6.5-cm square (Edmund Scientific Company, 101 East Gloucester Pike, Barrington, NJ 08007-1380, Phone: 609-573-6260)

Cardboard with 6- x 6-cm square cut in middle

Overhead projector

Cardboard large enough to cover the entire stage of the overhead projector

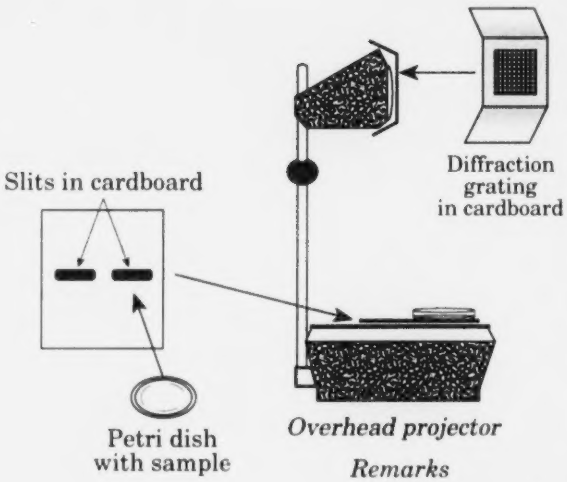
Safety

Be careful not to spill liquids on the overhead projector stage.

Procedure

Tape the 6.5- x 6.5-cm square of diffraction grating in the center of the piece of cardboard with a 6- x 6-cm square cut out of it. This grating will be placed over the upper projection lens of an overhead projector. Using the piece of cardboard large enough to cover the entire stage of an overhead projector, cut two rectangular "slits" (1- x 8-cm). The slits should be cut so that, when the cardboard is placed on the overhead stage, one slit is above the other (see illustration).

When the room is darkened, this arrangement will produce two continuous spectra on the screen. Over one of the slits, place a Petri dish. Pour a very dilute solution of potassium permanganate into the Petri dish. This "slit" will now produce a spectrum created by the presence of the compound in the Petri dish. This demonstration can be repeated for two other colored compounds (food coloring can also be used), showing that the bands produced are characteristic for the colors of these compounds.



Remarks

This demonstration will require practice to establish proper lighting in the room, arrangement for best projection of the spectra, and concentration of solutions that will produce the best results.

TIPS FOR THE TEACHER

Language of Chemistry

coordination complex compound or complex ion in which a metal ion or atom is bonded to one or more ligands.

filtrate liquid that has passed through a filter.

ligand molecule or ion bonded to the central metal atom in a coordination complex.

precipitate insoluble substance formed from soluble reactants.

qualitative analysis chemical analysis used to determine the identity of a sample.

quantitative analysis chemical analysis used to determine the amount of the constituents of a sample.

residue solid left after a liquid is evaporated or removed.

spectrometer instrument used to observe the absorption or emission of particular frequencies of light from a sample.

supernatant liquid solution remaining after a precipitate is formed.

Common Student Misconceptions

1. **"Insoluble means nothing dissolves."**

   Solubility is a relative term defined quantitatively by a solubility product (an equilibrium constant). At the qualitative level, an insoluble substance is defined in many texts as a substance that can only form solutions of less than 0.01 M.

2. **"Separations are clean."**

   No technique can completely and perfectly separate the components of a mixture.

3. **"Analysis is easy."**

   Chemical analysis is sometimes expected to answer such questions as: "What's in this barbecue sauce?" It is thought that all that is necessary is to perform a simple test or analyze a sample with an instrument, and the answer reveals itself. Actually, quite a lot of intellectual effort is often necessary.

History: On the Human Side

See "The Past 100 Years in Analytical Chemistry," by G. Ewing in *References*.

Two anecdotes (Oesper, 1975) are related to flame tests and the qualitative identification of metal ions. The first of these concerns Frederick Soddy (1887-1956). During a visit to Paris, he received a few milligrams of radium bromide from Marie Curie. When he returned to London and the laboratory of William Ramsay, for whom he worked, he watched in horror as Ramsay used up most of his sample to do an unnecessary flame test to confirm that the sample contained radium. The second anecdote concerns Robert Wood (1868-1955), who studied chemistry under Ira Remsen at Johns Hopkins University. Wood took his meals at a boarding house. It seems that whenever steak was served for supper, beef hash was served at the following breakfast. The patrons of the boarding house suspected that table scraps were used in preparing the hash. One evening Wood left several pieces of steak on his plate after he sprinkled on it what looked like table salt, but was actually lithium chloride. The next morning he burned a piece of hash in the table lamp flame, saw the brilliant red color characteristic of lithium, and proved to the delight of his fellow diners that their suspicions were correct.

Humor: On the Fun Side

1. Sing a song of Sulfide
   A beaker full of lime...
   Four and twenty test-tubes
   Breaking all the time.

   When the cork is taken out
   Fumes begin to reek...
   Isn't that an awful mess
   To have five times a week?

   CHEM 13 NEWS, October 1969, p. 84

2. **TEN COMMANDMENTS FOR STUDENTS OF ANALYTICAL CHEMISTRY**
   1. Thou shalt honor thy neighbor's olfactory sensibilities.
   2. Thou shalt not kill unless thou canst prove that he did spit in thy unknowns.
   3. Thou shalt not pencil titrate.
   4. Thou shalt not take the name of thy teacher in vain.
   5. Thou shalt not remain forever on the pans of thy balance.
   6. Thou shalt keep thy desk spotless that thy days may be long in this laboratory.
   7. Thou shalt not commit adulteration—of reagents.
   8. Thou shalt love the lab assistant as thyself.
   9. Thou shalt not covet thy neighbor's success, nor clean equipment, nor anything that is thy neighbor's.
   10. Two days shalt thou labor, sweat and swear in thy laboratory; the other days are set aside for thy teacher's peace of mind and for the conservation of thine own sanity.

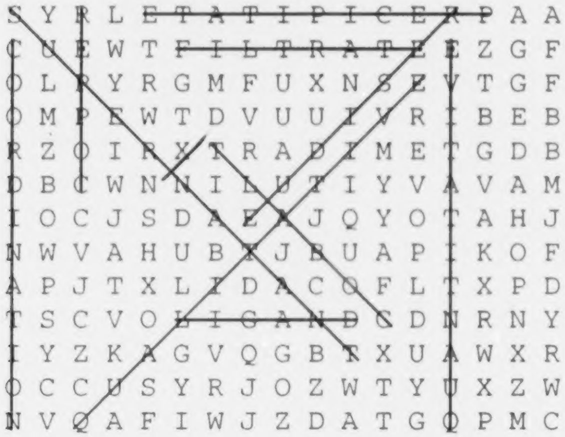
   CHEM 13 NEWS, January 1978, p. 1292

3. **LITTLE ION**

Little Ion in my flask
Do you mind much if I ask
What your name is little ion?
Can't you see you've got me cryin'?
Are you silver, lead or tin?
Don't you see me growing weaker
As I gaze into my beaker?
Can you bear to see my fate
If my unknown is turned in late?
Come now do not be a fink,
Are you cobalt? Iron? Zinc?
Stir and centrifuge and test
I have done my very best.
Yet my prof will probably say—
Chloride, nickel, sulfate, nay.
Little unknown be a doll
Help me pass the course called "Qual."

Courtesy of Jerry Wellins, Arthur Breyer, and Marie Bonner.

4. Word Search (see *Appendix* for master copy)



S	Y	R	L	E	T	A	T	I	P	I	C	E	R	P	A	A
C	U	E	W	T	F	I	L	T	R	A	T	E	E	Z	G	F
O	L	R	Y	R	G	M	F	U	X	N	S	E	V	T	G	F
O	M	P	E	W	T	D	V	U	U	I	V	R	I	B	E	B
R	Z	O	I	R	X	T	R	A	D	I	M	E	T	G	D	B
D	B	C	W	N	N	I	L	U	T	I	Y	V	A	V	A	M
I	O	C	J	S	D	A	E	X	J	Q	Y	O	T	A	H	J
N	W	V	A	H	U	B	T	J	B	U	A	P	I	K	O	F
A	P	J	T	X	L	I	D	A	C	O	F	L	T	X	P	D
T	S	C	V	O	L	I	G	A	N	D	S	D	N	R	N	Y
I	Y	Z	K	A	G	V	Q	G	B	T	X	U	A	W	X	R
O	C	C	U	S	Y	R	J	O	Z	W	T	Y	U	X	Z	W
N	V	Q	A	F	I	W	J	Z	D	A	T	G	Q	P	M	C

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Solid left after a liquid is evaporated or removed.
2. Type of chemical analysis used to determine a sample's identity.
3. Adjective describing the liquid remaining after a precipitate is formed.
4. Liquid that has passed through a filter.
5. Type of chemical analysis used to determine the amount of the sample's constituents.
6. Insoluble substance formed in an aqueous solution from soluble reactants.
7. Type of complex in which a metal ion is bonded to one or more ligands.
8. Molecule or ion bonded to a central metal atom.
9. The ion studied in the activity that forms a red precipitate with $Fe(CN)_6^{3-}$.
10. The ion studied in the activity that forms a dark blue solution with NH_3.

Answers: 1. RESIDUE 2. QUALITATIVE 3. SUPERNATANT 4. FILTRATE 5. QUANTITATIVE 6. PRECIPITATE 7. COORDINATION 8. LIGAND 9. COBALT 10. COPPER

5. See cartoons at end of module.

MEDIA

1. *Molecular Spectroscopy*, CHEM Study film/video available from Ward's Natural Science Establishment, Inc., P.O. Box 92912, Rochester, NY 14692-9012; (800) 962-2660.
2. Software published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
   - a. *Organic Spectroscopy*. Vol. II A, No. 2, for Apple II computer. Simulated IR, Proton and C-13 NMR, and mass spectra for a variety of compounds.
   - b. *HPLC – An Instrument Simulator*, Vol. I B, No. 2, for IBM PS/2 PC-compatible computers.
   - c. *Proton NMR Spectrum Simulator*, Vol. II C, No. 1, for Macintosh computer.
3. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).
   - a. For the Apple Macintosh computer: MC 901
   - b. For the Apple II computer: AP 901, AP 903, AP 904
   - c. For IBM PCs and PC-compatibles: PC 3401

   Six-Solution Problem – A simulation of the problem of determining the composition of six solutions by mixing them on a spot plate.

   Canal 1,2,3,4,5 – Tests a student's knowledge of the traditional qualitative analysis scheme.

EQUIPMENT AND INSTRUMENTATION

The Spectronic 20 is an instrument used to characterize solutions that can be useful in qualitative analysis. This type of instrument is sometimes found in high school laboratories.

Links/Connections

WITHIN CHEMISTRY

1. Within chemistry, the most obvious connections are with chemical equilibrium and acid/base chemistry. The separations of most inorganic qualitative analysis schemes depend on either the relative solubility of the metal ions with various anions or upon the ability of these ions to form complex ions. pH very often plays a role in these processes. The ability of H^+ to dissolve insoluble hydroxides, and for amphoteric hydroxides [*e.g.*, $Al(OH)_3$] to dissolve in excess base depends on competing equilibria. The solubility of other substances in ammonia depends on the formation of stable complex ions, another equilibrium consideration. There are also links to the *Transition Elements; Rocks, Minerals, and Gems; and Photochemistry* modules. The *Industrial Inorganic Chemistry* module has an exercise on the analysis of fertilizers.
2. **Alloys**. The purpose of preparing an alloy is to improve the specific properties of the principal component. Properties such as hardness, melting point, *etc.* can be controlled by alloying.
    a. Most gold jewelry is really a gold alloy containing gold, copper, and silver, since pure gold is too soft for ordinary use. 24 carat gold is 100% gold. 12 carat gold is 50% gold, 35% copper, and 15% silver. Coinage gold is usually 90% gold and 10% copper. White golds are mixtures of gold with either palladium or platinum. German silver and nickel silver contain copper, nickel, and zinc in different proportions.
    b. Copper is the principal component of brasses and bronzes. Red brass is 90% copper and 10% zinc, yellow brass is 67% copper and 37% zinc. Gun-metal bronze is 90% copper and 10% tin; bearing bronze is 82% copper, 16% tin, and 2% zinc; hardware bronze is 89% copper, 9% zinc, and 2% antimony.
    c. Sodium melts at 97.8 °C, potassium melts at 63.7 °C, but an alloy of 22% sodium and 78% potassium melts at –11 °C. Soft (lead) solder is composed of tin, lead, antimony, copper, and bismuth, whereas hard (silver) solder contains 10-80% silver and varying amounts of copper and zinc.

TO THE CONTEMPORARY WORLD

See *Forensic Chemistry* module.

Water purity.

Environmental hazards (Pb, NO, sulfur in industry).

Community resources: water treatment plant, soil conservation (county agent), sewage treatment facility, forensic facility.

Technicians use qualitative testing methods to test pool water for chlorine and acid and indicators to decide if the pH is too high or too low.

Extensions

Chromatography is an analytical technique that is easily done in a high school laboratory. *Demonstration 3* in the *Separations* and *Solutions* modules illustrates paper chromatography. *Demonstration 4* in the *Solutions* module describes a column chromatography activity. (See also Dreyfus Microscale Lab: Paper Chromatography, Brogie, 1987.)

References

Module drafted by Mike Abraham, Donna Deen-Coshow, and Bill Fix, the Oklahoma team.

Abraham, M., and Pavelich, M. (1991). *Inquiries into chemistry* (2nd Ed.). Prospect Heights, IL: Waveland Press.

"Chemical Tests," "Laboratory Procedures," and "Qualitative Ion Analysis Using Semi-Micro Techniques" are taken from this work.

Barber, J. (1985). *Crime lab chemistry* [LHS GEMS] (Teacher's Guide). Berkeley, CA: Lawrence Hall of Science, University of California.

From the Great Explorations in Math and Science (GEMS) series, this contains activities relating to chromatography, separating mixtures, pigments, and solubility. Although written for Grades 4-8, the activities are also suitable for high school students. Order this book from Lawrence Hall of Science, University of California, Berkeley, CA 94720.

Brogie, E. (1987). *Dreyfus microscale lab: Paper chromatography* [Woodrow Wilson-Dreyfus Institute, 1987, #6]. Princeton, NJ: Princeton University.

Chen, C., and Ingels, R. (1987). *Dreyfus microscale lab: Six solutions* [Woodrow Wilson-Dreyfus Institute, 1987, #35]. Princeton, NJ: Princeton University. Published by University of Nebraska in *Microscale Experiments for High School Chemistry Classes,* 1988. [Write to David Brooks, 118 Henzlik Hall, University of Nebraska, Lincoln, NE 68588-0355.]

Chen, C., and Ingels, R. (1987). *Dreyfus microscale lab: Qualitative analysis* [Woodrow Wilson-Dreyfus Institute, 1987, #36]. Princeton, NJ: Princeton University.

Cole, G. M., and Waggoner, W. H. (1983). Qualitative analysis: The current status. *Journal of Chemical Education, 60,* 135-140.

The current status according to these authors is that qualitative analysis is making a comeback. The article is a reconsideration of the qual scheme to make it more suitable for modern times.

Elo, H. (1987). Is thioacetamide a serious health hazard in inorganic chemistry laboratories? *Journal of Chemical Education, 64,* A144, A146.

The short answer from this article is "yes."

Ewing, G. W. (1976, April 6). The past 100 years in analytical chemistry. *Chemical and Engineering News, 54*(15), 128-145.

This is a good source of history for this topic.

Gerber, S. M. (Ed.). (1983). *Chemistry and crime.* Washington, DC: American Chemical Society.

Numerous fictional and nonfictional references to chemistry and its relationship to crime detection.

Lagowski, J. J., and Sorum, C.H. (1991). *Introduction to semimicro qualitative analysis* (7th Ed.). Englewood Cliffs, NJ: Prentice-Hall.

This is the latest edition of a widely used text on inorganic qualitative analysis. It is cited here only as an example. There is a large number of such texts available and any one of them is a useful reference not only for qualitative analysis but for general chemical information.

Oesper, R. E. (1975). *The human side of scientists.* Cincinnati, OH: University of Cincinnati Publications.

This is the source of the two anecdotes that are related to flame tests and the qualitative identification of metal ions (in *History* section).

Rich, R. L. (1984). Qualitative analysis, with periodicity, for "real" solutions. *Journal of Chemical Education, 61*, 53-58.

This article describes two qualitative analysis schemes based on bromide and iodide ions. It is claimed that these schemes make more "sense" in terms of periodicity.

Solomon, S., Fülep-Poszmik, A., and Lee, A. (1991). Qualitative analysis of eleven household compounds. *Journal of Chemical Education, 68*(4), 328-329.

This article contains a scheme for identifying compounds commonly found in the house. All of the chemicals used, including reagents, can be obtained as consumer items.

West, P. W., and Vick, M. M. (1959). *Qualitative analysis and analytical chemical separation* (2nd Ed.). New York, NY: MacMillan.

This is a qualitative analysis scheme based on hydroxide and benzoate ions instead of sulfide.

Appendix

- **Transparency Masters**
  1. Identification of Selected Substances
  2. Qualitative Ion Analysis Using Semimicro Techniques
  3. Word Search
- **Humor**

Identification of Selected Substances

Products of chemical reactions can be identified by their chemical or physical properties. Following is a key to identifying some common substances.

Formula	Color	Odor	State at 25 °C	Solubility in H_2O	Miscellaneous
H_2	None	None	Gas	No	Burns explosively
O_2	None	None	Gas	No	Makes glowing splint flame
CO_2	None	None	Gas	Yes	Turns limewater milky
NO_2	Brown	Choking	Gas	Yes	Solution is acidic
I_2	Purple	Medicine-like	Solid	Yes	Purple in hexane (or cyclohexane)
Br_2	Red-brown	Choking	Liquid	Yes	Orange in hexane (or cyclohexane)
Cl_2	Pale green	Choking	Gas	Yes	Yellow in hexane (or cyclohexane)
NH_3	None	Choking (smelling salts)	Gas	Yes	Solution is basic
SO_2	None	Suffocating	Gas	Yes	Solution is acidic
H_2S	None	Rotten eggs	Gas	Yes	Solution is neutral
HCl	None	Choking	Gas	Yes	Solution is acidic
H_2O	None	None	Liquid	________	$CoCl_2$ paper turns pink

Qualitative Ion Analysis Using Semimicro Techniques

In these experiments you will be concerned with determining what particular ions are in a solution or solid. This concern with only what is present is called qualitative analysis. (A concern for how much is present is called quantitative analysis). You will be working with very small amounts of solutions, usually about 1/4 to 1/2 of a milliliter. Thus the experimental procedures are called semimicro (small amount) techniques. Two other definitions of importance are:

Precipitate a solid formed when two solutions are mixed (abbreviated ppt).

Supernatant liquid the liquid phase of a solid plus liquid mixture (abbreviated spn).

The basic idea in these experiments is to add a reagent to a solution of ions and to note any and all changes that occur. In particular, one looks for and notes down:

1. The formation and color of a ppt.
2. Changes in color of the solution or spn.
3. Redissolving of a ppt.
4. Evolution of a gas.

Additions and Mixing

The amounts of solutions to be mixed in each test are often critical to the results and must be carefully watched. The amounts are usually controlled by counting drops of solution used (18 drops = 1 mL). When making additions suspend the end of the dropper just above the opening of the test-tube and gently apply pressure to the dropper bulb so that the drops fall one at a time into the test-tube.

When test reagent has been added to a solution, mix the two completely before you write down your observations. The results that you see when one solution is layered on top of the other may not be the same as when they are mixed. To mix the solutions move a clean stirring rod gently up and down within the test-tube; *i.e.*, use a "butter churn" type motion. Watch that the liquid does not overflow the test-tube. Be careful not to "punch" a hole in the bottom of the test-tube.

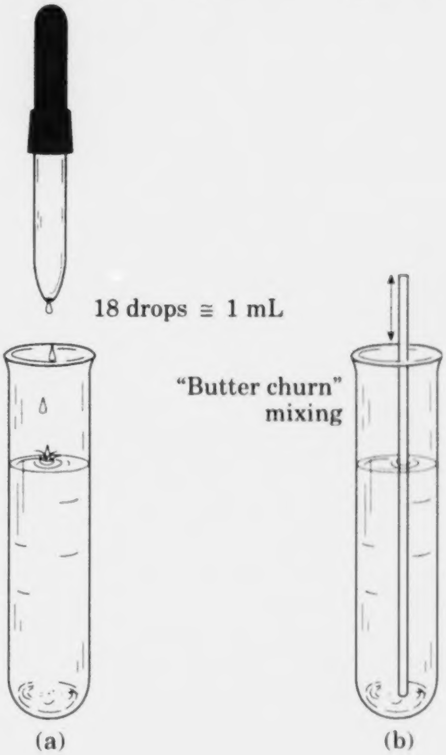


Figure A. (a) Dropwise addition and (b) mixing.

The Centrifuge

The centrifuge is a machine that spins a test-tube solution at high speeds causing the heavier and lighter parts of the mixture to separate. In essence, it accelerates what gravity would have done anyway. The centrifuge is used to separate a precipitate from the supernatant liquid. Usually only 1 or 2 min of centrifuging are needed to separate the ppt and spn.

Precautions to be observed in using the centrifuge are as follows. (1) Do not have the test-tubes more than 4/5 full. This will prevent spilling in the centrifuge. (2) Place the tubes in the centrifuge in a symmetrical pattern such as one tube directly across from another. If you are centrifuging just one test-tube, place another containing an equal volume of water across from it. If the placement of the tubes is not balanced, the spinning will be uneven and the centrifuge will move ("walk") on the bench. Allow the centrifuge to slow down by itself. Helping it slow down can cause injury. Also if the slow down is too abrupt the ppt and spn will remix.

Decanting

This is the term used to denote removing a supernatant liquid from a precipitate. There are two decanting methods. The first is to simply pour the spn out of the test-tube leaving the ppt behind. The second method is to draw the spn out of the test-tube using a dropper. This dropper method is used when the ppt is not a firm pack and might pour out with the spn. It is also the preferred decantation method where the spn is to be saved for later use. It is often difficult to pour directly from one small test-tube to another.

When using the dropper method be sure to compress the bulb **before** inserting the dropper. A rush of bubbles in the spn will cause the ppt and spn to remix.

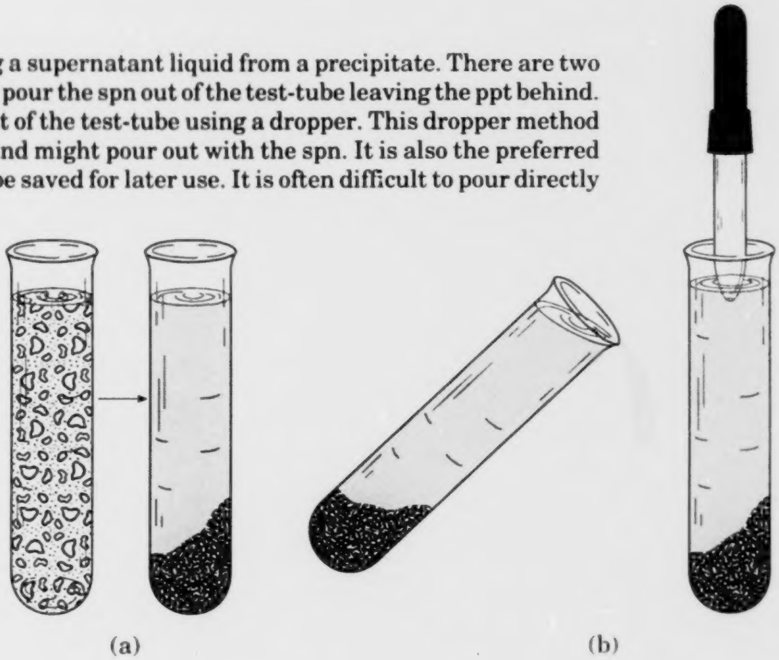


Figure B. (a) Effect of centrifuging and (b) decanting by pouring or drawing.

Washing a Precipitate

Not all of the supernatant is removed from the test-tube when one decants. A few drops remain on top of the precipitate, some clings to the test-tube walls and some is trapped within the packed precipitate. If a later test on the ppt requires that the chemicals in the spn be absent, these traces of spn are removed by a procedure called washing.

Washing a ppt involves the following sequence of steps. (1) Decant the original spn. (2) Add to the ppt about 1 mL of distilled water (or other prescribed liquid). (3) Completely mix the H_2O and ppt. This allows the H_2O to leach out any spn trapped in the packed ppt. (4) Centrifuge the mixture. (5) Decant the liquid, discarding it. (6) Repeat the process. Usually two such water washes are sufficient as they will leach out about 97% of the spn originally left in the test-tube.

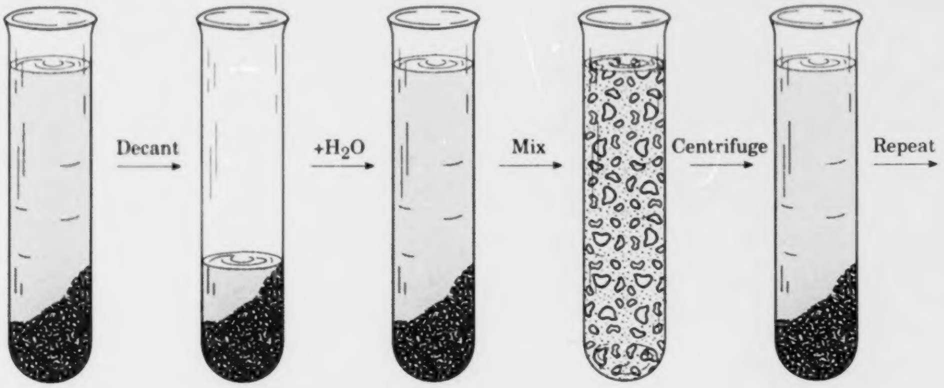


Figure C. Washing a precipitate.

Heating

You should **not** use a burner flame to directly heat liquid in a small test-tube. Such direct heating is too intense and too localized. It will usually cause the liquid to blow out of the test-tube. This practice can cause injury and, in any case, ruins the experiment.

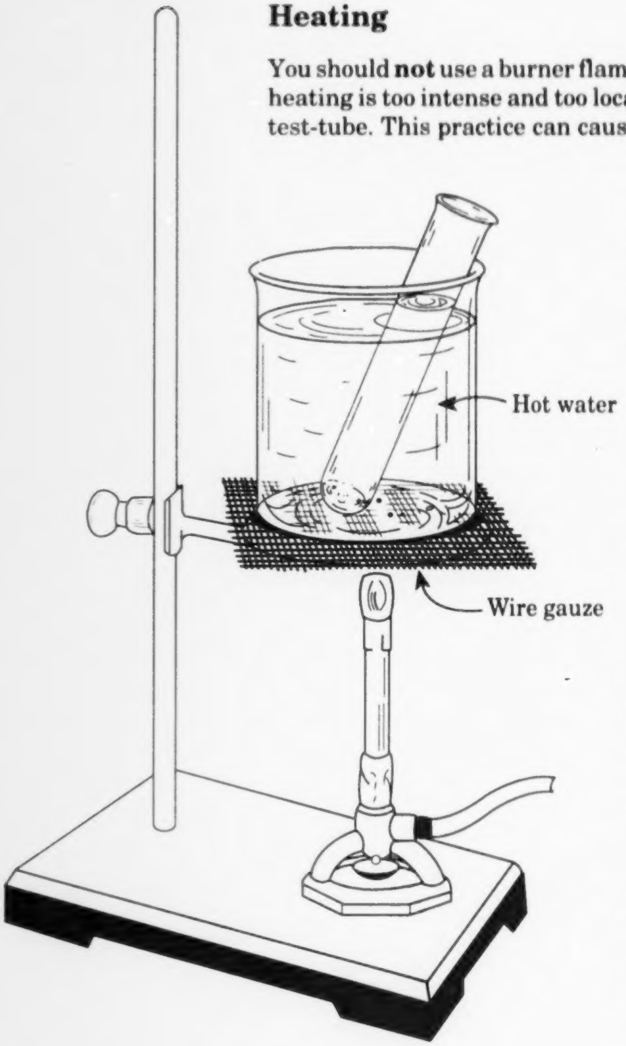


Figure D. Heating a sample.

Heating is to be accomplished by resting the test-tubes in beakers of heated water. This heating is more even and keeps the temperature from exceeding 100 °C. If a test-tube heating platform (also called an aluminum holder) is available it should be used in the beaker. Otherwise use a beaker large enough to keep the test-tubes from protruding too far over the beaker lip.

Adjust the water level in the beaker so that only the lower 1/2 to 3/4 of the test-tube is submerged. Move the burner flame in and out as needed to keep the water hot but to prevent violent boiling (see Figure D).

pH Measurement and Control

It is important in many reactions that the solution's pH be carefully controlled. The control is achieved by adding an acid or base solution dropwise until the required pH is obtained. The pH of the solution should be measured after each dropwise addition by using litmus or pH paper. Do not dip the pH paper into the solution. This practice can lead to chemicals from the paper getting into the solution. The proper procedure is to mix the solution with a stirring rod and then to touch the wet end of the rod to the test paper.

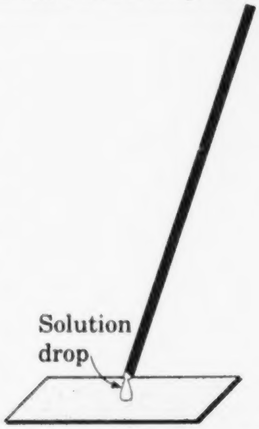


Figure E. Testing pH.

The pH of a gas bubbling out of a solution can be tested by wetting the test paper with water and then holding it over the mouth of the test-tube.

Contamination

Even the smallest contamination of your ion or reagent solutions can change the test results. Contamination is a particular problem with these experiments because you reuse equipment so often. **Clean and distilled water rinse each piece of equipment after each use.** Thus you should clean a dropper before using it in a different solution, clean a stirring rod before using it to mix a second solution, *etc.*

Often cleaning requires only that the test-tube, rod, *etc.* be rinsed several times with tap water then 2 to 3 times with distilled H_2O. The equipment does not have to be dry. Use a test-tube brush to clean out stubborn ppts. A dropper can be cleaned by removing the bulb and running water down the barrel.

Complete Precipitation

In analysis of an unknown you will often depend on precipitating an ion completely from solution to separate it from the other ions. It is important that this precipitation reaction be complete.

After centrifuging the ppt plus spn mixture, add a drop or two more of the test reagent. If new ppt forms in the spn, add 2 more drops of test reagent, mix, and recentrifuge. Again add a drop of reagent. Continue this 1 to 2 drop testing of the spn until no new ppt forms. At this point the reaction is complete (see Figure F).

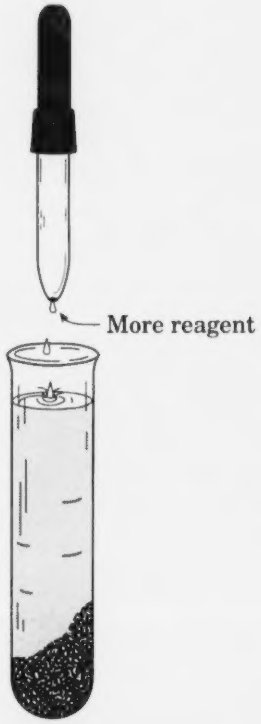


Figure F. Testing for complete precipitation.

Organization and Reagents

It is strongly suggested that you take a few minutes to plan your work. You will be performing 25 or more individual tests in one period; thus, some organization is needed to save time and not to confuse your results. The following ideas might be helpful.

1. After using a dropper, immediately return it to its reagent bottle. Do not have two droppers out of their reagent bottles at one time. This will keep you from returning a dropper to the wrong reagent bottle.
2. Label each test-tube and assign it a labeled position in your test-tube rack. Never have two test-tubes out of the rack at the same time; always return one before picking up another. This keeping track of test-tubes is needed because the label on the tube can be smeared in handling. Be sure the label on a tube is clear before you put it into a centrifuge.
3. Put all used, dirty glassware in one section of the bench so as not to confuse it with clean glassware.
4. When you are to run a test on several different ion solutions, run the test on all of them simultaneously. Set up a labeled tube for each ion solution in the rack. Put the proper ion solution into each test-tube. Next go from one to the other adding the test reagent and mixing. This simultaneous testing saves time and allows you to make direct comparisons of the results for different ion solutions.
5. Both lab partners should observe the results of each test and record their own perception of the results. Also describe fully your observation; do not simply put down "Reacted." You will need to recall exactly what you saw happen when you analyze your unknown.

Word Search

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S Y R L E T A T I P I C E R P A A
C U E W T F I L T R A T E E Z G F
O L P Y R G M F U X N S E V T G F
O M P E W T D V U U I V R I B E B
R Z O I R X T R A D I M E T G D B
D B C W N N I L U T I Y V A V A M
I O C J S D A E A J Q Y O T A H J
N W V A H U B T J B U A P I K O F
A P J T X L I D A C O F L T X P D
T S C V O L I G A N D C D N R N Y
I Y Z K A G V Q G B T X U A W X R
O C C U S Y R J O Z W T Y U X Z W
N V Q A F I W J Z D A T G Q P M C
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Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Solid left after a liquid is evaporated or removed.
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7. Type of complex in which a metal ion is bonded to one or more ligands.
8. Molecule or ion bonded to a central metal atom.
9. The ion studied in the activity that forms a red precipitate with $Fe(CN)_6^{3-}$.
10. The ion studied in the activity that forms a dark blue solution with NH_3.

"We're going to have to switch to semi-micro. This is ridiculous!"

CHEM 13 NEWS, December 1980, p. 3. Reprinted with permission.

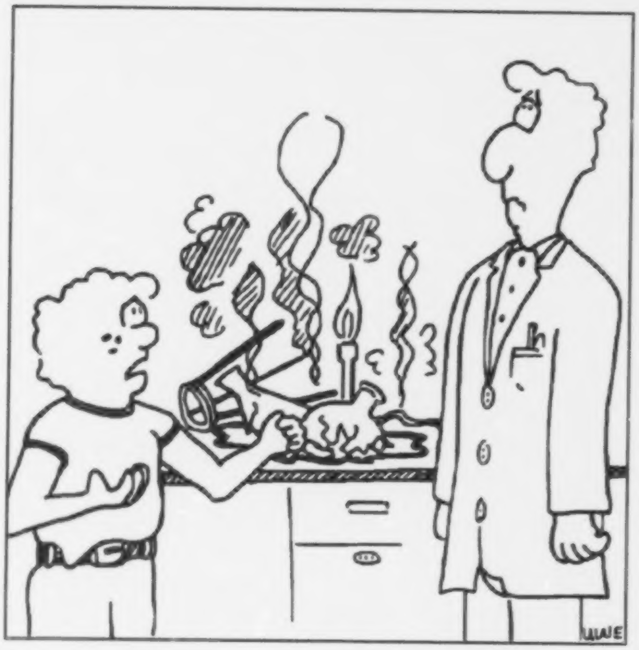
LET GO BROWN !! LAB'S OVER !!
LWE

CHEM 13 NEWS, October 1978. Reprinted with permission.

357

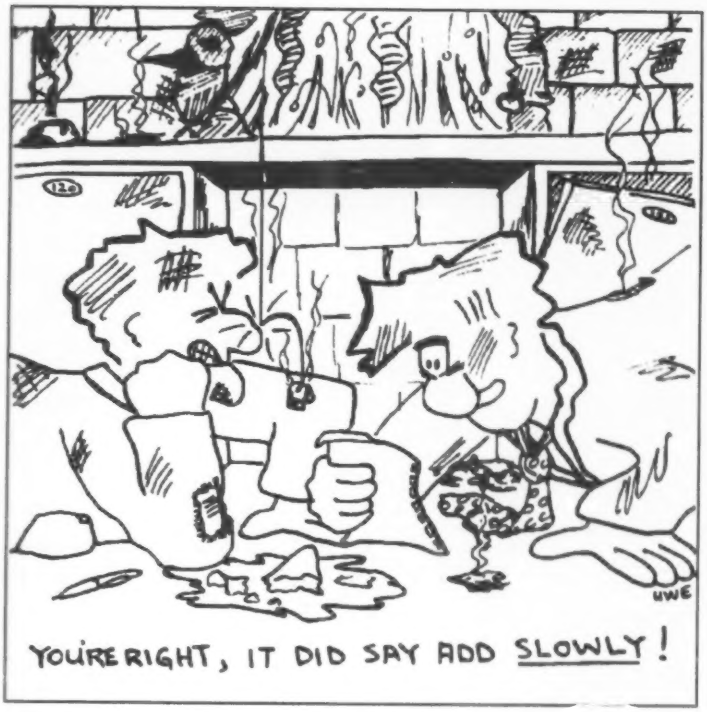
UWE
YOUR LAB TECHNIQUE ALWAYS STANK GEORGE!!!

CHEM 13 NEWS, November 1978. Reprinted with permission.



IT MUST HAVE BEEN CRACKED WHEN I GOT IT.

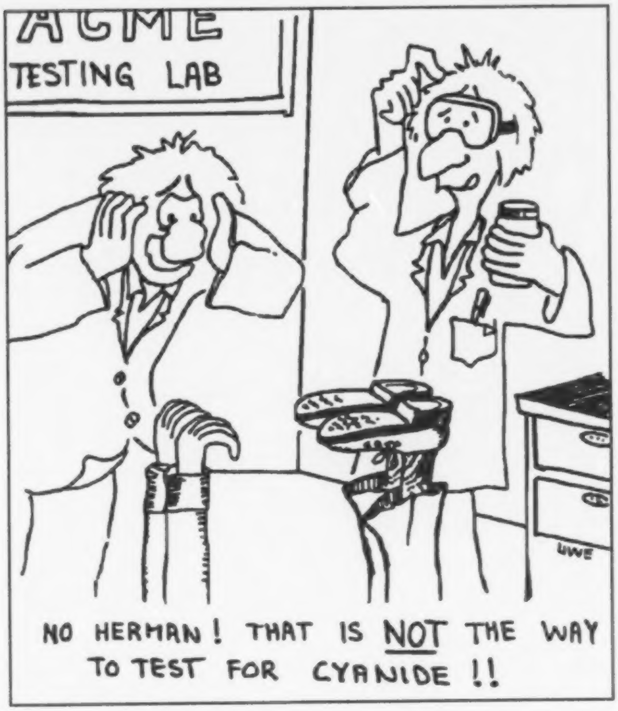
CHEM 13 NEWS, March 1980, p. 6. Reprinted with permission.


12a
UWE
YOU'RE RIGHT, IT DID SAY ADD SLOWLY!

CHEM 13 NEWS, March 1981, p. 9. Reprinted with permission.

OH WOW! PORSCHE LAB GOGGLES!
Hobbs

CHEM 13 NEWS, September 1981, p. 5. Reprinted with permission.


ACME
TESTING LAB
UWE
NO HERMAN! THAT IS NOT THE WAY TO TEST FOR CYANIDE!!

CHEM 13 NEWS, February 1983, p. 13. Reprinted with permission.

A SourceBook Module

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RATES OF REACTION

ChemSource

Instructional Resources for Preservice and
Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

A rocket leaving earth's gravitational field may travel at 40,000 km/h. A race car might speed along at 400 km/h, while a snail may crawl along at 0.00004 km/h. Each object travels at its own rate of speed. Chemical reactions behave similarly.

Dynamite explodes. Iron rusts. One chemical reaction is over in an instant; the other proceeds almost imperceptibly slowly. The difference lies in the rates of the reactions. Rates of chemical reactions can be measured experimentally. Methods include extent of color change, conductivity change, gas volume change with time, *etc.*

The rate at which a chemical reaction occurs depends on many factors, including temperature, concentration, surface area and the addition of a catalyst.

Have you ever wondered why campers use kindling wood to light their campfire, instead of lighting the large logs directly with matches? (Surface area) Or why fruit spoils more quickly in an open fruit bowl than it does when stored in the refrigerator? (Temperature) Or why acetylene burns with a blue flame, hot enough to weld metals together, if given enough oxygen, as in an oxy-acetylene torch; yet burns with a yellow, sooty flame in "regular" air? (Concentration) Or why wet firewood is so much harder to light than dry kindling? Answers to these and many other interesting and puzzling questions can be found through a study of reaction rates.

Chemists explain rates of chemical reactions by means of *kinetic molecular theory* (collision theory), which states that molecules and/or atoms must move and collide in order to react. Not only must they collide, but they must collide forcefully enough and with proper geometry (spatial arrangement) to break old chemical bonds and form new ones. Everyday phenomena such as those listed earlier can be readily explained using collision theory.

Reaction rates play a significant role in everyday life. Refrigerators and freezers reduce temperatures and thus chemical reaction rates within microorganisms responsible for food spoilage. Inhibitors also are used to reduce biochemical reaction rates in order to minimize food spoilage. In the chemical industry, control of reaction rates frequently determines product yield and the extent of profit realized. Catalytic converters on automobiles increase the rate of combustion of various hydrocarbon pollutants to carbon dioxide and water vapor, and the rate of reduction of nitrogen oxide pollutants to nitrogen gas. Research continues in efforts to retard the aging process, composed of a complex series of biochemical reactions. Reaction rates or reaction kinetics is a vital and vigorous area of study.

Linus Pauling defines chemistry as "the science of substances, their structures, their properties and the reactions through which they are converted to new substances." These reactions involve definite quantities of matter and energy and require a specific set of experimental conditions. Under these conditions the reactions proceed at definite rates to give definite yields of product(s). In order for a synthesis reaction to be safe and of practical value, its rate must not be too rapid or so slow that the product cannot be made economically. An understanding of the mechanism of a reaction better enables chemists to control the rate of the reaction.

PLACE IN THE CURRICULUM

A study of rates of reaction normally follows the discussion of chemical equilibrium. Students should be aware of the fact that a reaction can move forward or in the reverse direction, that there are some critical conditions under which the reaction does move forward, and that these conditions also determine the rate of the reaction. Since many of the examples usually used to illustrate reaction rates are for substances in the gaseous state, it is helpful to discuss gases and review the gas laws prior to studying reaction rates.

CENTRAL CONCEPTS

1. Chemical reaction rates can be determined experimentally by measuring the change in concentration of a reactant or product per unit of time.
2. Reaction rates are influenced by factors such as nature of the reactants, presence of a suitable catalyst, temperature, concentration of the reactants, and available surface area of solid or liquid reactants, or pressure for gases.
3. Molecular-scale models known as *reaction mechanisms* are used to explain a reaction rate's dependence on the factors listed in Concept 2.
    a. *Collision theory* assumes reactant molecules must hit each other hard enough, and in the proper orientation, to be rearranged into new (product) molecules if a reaction is to occur. These are known as effective collisions.
    b. A transitory *activated complex* is assumed to form in each effective collision, which then breaks apart into new (product) species.
    c. A reaction's balanced equation rarely implies the rate equation for that reaction, and never indicates the rate of reaction. In part, this is because the balanced equation rarely represents the reaction mechanism—a sequence of steps by which the reaction is assumed to proceed.
4. A potential energy diagram may be constructed for each reaction, which may be thought of as a side-view "road map" of the reactants' journey to form products.
    a. The potential energy barrier to a reaction is represented as a "hump" within a potential energy diagram. The energy needed to surmount the potential energy barrier is known as *activation energy.*
    b. A *catalyst* is assumed to provide an alternate reaction path requiring lower activation energy.
    c. *Heat*, or *enthalpy* change of a reaction, can be represented as the difference in energy levels between reactants and products on a potential energy diagram.

RELATED CONCEPTS

1. Chemical equations
2. Mole concept—stoichiometry
3. Solution concentration—molarity
4. Kinetic molecular theory
5. Partial pressures
6. Molecular structure and shape
7. Thermochemistry—free energy of reaction
8. Equilibrium (Since $K_{eq} = k_f/k_b$ for each step of a reaction mechanism, the equilibrium constant, K_{eq}, is explicitly related to the rate constants k_f and k_b.)

RELATED SKILLS

1. Use of pipets, burets, graduated cylinders, and timers
2. Problem solving techniques
3. Exponential notation
4. Dimensional analysis
5. Estimating answers
6. Unit conversions

PERFORMANCE OBJECTIVES

Upon completion of their study of rates of reactions, students will be able to:

1. define kinetic molecular theory (collision theory) and use it to explain why chemical reactions occur.
2. define kinetic energy in terms of velocity (or speed) and mass of particles.
3. relate molecular motion and energy transfer to temperature.
4. compare velocities and masses of different gas molecules, measured at the same temperature.
5. describe three common states of matter in terms of relative kinetic and potential energy.
6. describe factors that can affect reaction rate.
7. explain the meaning of "reaction mechanism."
8. distinguish a reaction mechanism from the overall equation describing a reaction.
9. draw and label the key parts of a kinetic energy distribution curve.
10. define threshold energy.
11. draw and label the key parts of a potential energy diagram.
12. define activation energy.
13. describe how a catalyst is able to affect the rate of a chemical reaction.
14. predict the effect of changes in various experimental reaction conditions on the rate of a reaction.
15. distinguish between experimental changes that might increase the rate of a given reaction from those likely to decrease the rate.
16. discuss factors that would cause the rate of a specific chemical reaction to increase/decrease.
17. discuss reaction rates and the factors that change them in terms of simple collision theory.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: A Study of Reaction Rates

Introduction

Various factors affect reaction rates. Concentration and temperature, two such factors, will be explored in this laboratory activity. You will investigate a chemical reaction called a "clock reaction" to determine the effect of each of these two factors on reaction rate. *Part I* deals with concentration changes of one reactant. *Part II* involves temperature changes. You will carefully control other important variables as the activity is performed.

Clock reaction involves mixing of the two solutions. Solution A is a dilute solution of potassium iodate, KIO_3. This solution is the source of iodate ions, $IO_3^-(aq)$. Solution B contains soluble starch and the second reacting species, hydrogen sulfite ions, $HSO_3^-(aq)$.

The initial step of the reaction is represented by the equation

$$IO_3^-(aq) + 3HSO_3^-(aq) \rightarrow I^-(aq) + 3SO_4^{2-}(aq) + 3H^+(aq)$$

When hydrogen sulfite ions, $HSO_3^-(aq)$, are used up, iodide ions, $I^-(aq)$, react with remaining iodate ions, $IO_3^-(aq)$, to produce iodine, $I_2(s)$:

$$5I^-(aq) + 6H^+(aq) + IO_3^-(aq) \rightarrow 3I_2(s) + 3H_2O(l)$$

Molecular iodine (I_2) forms a blue substance—an iodine-starch complex—in the presence of starch from Solution B. The formation of blue color thus indicates that the reaction has proceeded to this point:

$$I_2(s) + \text{Starch}(aq) \rightarrow I_2\cdot\text{starch}$$

The iodine starch complex involves a ß-amylase-starch cage containing I_2, I^-, and H_2O in a three-dimensional structure inside the cage.

Purpose

To investigate concentration and temperature as two factors affecting the rate of chemical reactions.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Heating of beakers of water over an open flame is always potentially hazardous; burners should be handled with care.
3. Dispose of the used materials as your teacher directs.

Procedure

Part I

To investigate the effect of changing concentration of one reacting species on reaction time, you will prepare dilutions of Solution A, in which iodate ion, $IO_3^-(aq)$, concentration will vary. The concentration of the other reacting species, hydrogen sulfate ion, $HSO_3^-(aq)$, contained in Solution B, will be held constant. Since temperature change is not a consideration in this part of the laboratory activity, all solutions should be kept at room temperature throughout. Your teacher will assign each laboratory group several concentrations to investigate. By sharing data at the end of the activity, all laboratory groups will receive all the data. From these data you should be able to draw conclusions concerning the effect of concentration on reaction time.

1. Use a clean graduated cylinder to measure 10.0 mL Solution A; pour it into a clean test-tube (18- x 150-mm). Rinse the graduated cylinder. In a similar manner place 10.0 mL Solution B into another test-tube. If the solutions have been in the laboratory for some time you may assume that they are at room temperature. Otherwise, place the two test-tubes containing solutions into a 250-mL beaker about two-thirds full of water at room temperature and let them stand for several minutes.
2. Record the time to the nearest second as you pour Solution A into Solution B and then pour the mixture back and forth quickly three times to obtain uniform mixing. Alternately, stopper the test-tube and invert to mix. Time should be recorded from the first instant the two solutions are in contact.
3. Observe the solution in the tube carefully and record the time (to the nearest second) corresponding to the first sign of a reaction.
4. Repeat this procedure to check your results if so directed by your teacher.
5. Prepare different concentrations of the KIO_3 solution by diluting Solution A as specified below. Do as many dilutions as specified by your teacher.

mL Solution A	mL Distilled Water
9.0	1.0
8.0	2.0
7.0	3.0
etc.	*etc.*

   Note that the total volume is always 10.0 mL. Thoroughly mix each diluted solution (Solution A and distilled water) prior to allowing it to react with Solution B.

6. Repeat Steps 1-4 by adding one of the diluted solutions of KIO_3 to 10.0 mL Solution B, both at room temperature.

Part II

To investigate the effect of temperature on this reaction's rate you will determine the time for reaction at room temperature and at other temperatures within a range of 20 °C of room temperature. Your teacher will assign your laboratory group particular temperatures to investigate. By exchanging data with other groups, you will be able to draw conclusions regarding the effect of temperature on reaction time.

1. Pour 10.0 mL Solution A (labeled for *Part II*) into one test-tube (18 x 100 mm) and 10.0 mL Solution B into another. These solutions must be brought to the desired temperature before they are mixed. Put both tubes into a 250 mL beaker about two-thirds full of water at the temperature you were assigned to investigate. Let them stand for about 10 min so the solutions come to the temperature of the water bath.
2. Record the time to the nearest second as you pour Solution A into Solution B and then pour the mixture back and forth quickly three times to obtain uniform mixing. Time should be recorded from the instant both solutions are first in contact.
3. Place the test-tube back in the water bath. Observe carefully. Record the time (to the nearest second) corresponding to the first sign of a reaction.
4. Repeat the procedure at the same temperature to check your results if so directed by your teacher.

5. Repeat Steps 1-4 at another temperature, as specified by your teacher.
6. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

Part I

1. The concentration of potassium iodate in prepared Solution A is 0.020 M. Calculate the number of moles of KIO_3 contained in each milliliter of Solution A.
2. Calculate the initial molar concentration of KIO_3 in each mixture of A and B prepared in *Part I*.
3. Prepare a graph of your concentration-time data, with time on the vertical axis (ordinate) and concentration of KIO_3 on the horizontal axis (abscissa). Include the data of all other laboratory groups in your plot.

Part II

1. Prepare a graph of your temperature-time data, with temperature on the horizontal axis (abscissa) and time on the vertical axis (ordinate).

Implications and Applications

Part I

1. Why is it important to maintain the total volume at 10.0 mL during the dilutions of Solution A?
2. What generalizations can you make concerning the effect of varying the concentration on the time of the reaction?
3. How is the *time* of a reaction related to the *rate* of the reaction?

Part II

1. What general relationships can you derive from your temperature-time graph?
2. Predict the time of reaction at 0 °C and at 50 °C, assuming other variables in the laboratory procedure are kept constant.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: A Study of Reaction Rates

Major Chemical Concept

This exercise allows students to determine experimentally how various factors—notably concentration and temperature—affect the time for a reaction to occur, and thus how they affect the reaction rate. Emphasis is placed on quantitative results.

Level

All general students

Expected Student Background

Students should have a general idea of the concept of reaction rate. It would be helpful if they have seen examples of varying rates, at least on a qualitative basis.

Time

One or two class periods. The length of time available will determine how many different trials of concentration and temperature variations you will assign each student group.

Safety

Students should be cautioned about the hazards of an open burner flame at their bench.

Materials (For 24 students working in pairs)

Nonconsumables

24 Test-tubes, 18- x 150-mm
12 Beakers, 250-mL
12 Graduated cylinders, 10-mL
12-24 Thermometers, –10 to 110 °C
12 Clocks with sweep second hand, or digital watches with seconds
12 Burners
12 Ring stands and rings
12 Ceramic/wire gauze

Consumables

72-96 Ice cubes for cooling water bath
300 mL Distilled water at room temperature
1 L Solution A (4.3 g KIO_3/L; see *Advance Preparation* Note 4 for further dilution to be prepared)
1 L Solution B (0.2 g $Na_2S_2O_5$, 4 g soluble starch, 5 mL 1 M H_2SO_4/L) Note that $Na_2S_2O_5$ is sodium META-bisulfite, not sodium bisulfite (see *Advance Preparation* Notes 2 and 3)

Advance Preparation

1. To prepare each liter of Solution A, add 4.3 g of KIO_3 to a 1 L volumetric flask about half-full of distilled water. Stopper the flask and swirl until the solid dissolves. Fill to the 1 L mark with distilled water. Pour into suitable smaller containers for student use. Large burets work well for group dispensing, as do squeeze bottles such as laboratory wash bottles. If you choose this latter method, be certain the bottles are well-labeled to distinguish them from other bottles normally found in the laboratory.

2. For each liter of Solution B, make a paste of 4 g soluble starch in a small volume of water. Add this slowly to about 900 mL boiling water, and boil it for a few minutes. Allow the system to cool; then add the 0.2 g $Na_2S_2O_5$ and stir. This solution can be stored for a week or more prior to using it in this laboratory activity. (*Second Option:* Some teachers purchase commercial spray starch and use this to make the starch solutions very quickly by simply spraying the starch into a large volume of water. Try this method beforehand to determine how much spray starch is needed to give the expected deep blue color. After the starch solution has been prepared in this way, add the 0.2 g $Na_2S_2O_5$, sodium **meta**bisulfite, and stir.)
3. On the day of the activity, add the sulfuric acid to Solution B *just before use*. Stir the solution well and do a trial run with 10 mL A and 10 mL B. You may need to add more acid until the time is within the 10-15 second range. If the blue color appears faded, or if it forms gradually through the tube, you may need to add more $Na_2S_2O_5$. If the reaction occurs too slowly, at lower dilutions the reaction time will be too long for reasonable results. If the blue color appears too quickly, you can dilute Solution A. If the reaction occurs too rapidly (less than 10 s) at room temperature, reactions at higher temperatures will be too rapid to measure accurately, or they may not happen at all.
4. To prepare for *Part II*, dilute Solution A to about half its original concentration and test the room-temperature reaction time. It should be 20-25 s in order to prevent higher-temperature reactions from occurring too rapidly.
5. Be sure to pour back and forth three times in order to establish a standard mixing technique, ensuring more precise results. The technique should take only 5 or 6 seconds.

Pre-Laboratory Discussion

1. Avoid giving students any preconceived ideas about the effect of concentration on reaction rate (time). Previous student exposure has probably focused on qualitative observations rather than quantitative data.
2. A discussion of the steps in the reaction can be mentioned now, or can be dealt with as part of the post-laboratory activities. It may be sufficient to mention here that formation of blue color indicates presence of iodine molecules. Students may recall the iodine test for starch in previous biology classes.
3. You may find it helpful to demonstrate the reaction to students so that they will know what to expect when they do the activity themselves. This may prevent the loss of first-round results as students become amazed and miss the time of the color change when it finally appears (the "ooh and aah!" factor).

Part I

4. Assign all groups to do the same concentration first to help them understand the laboratory procedure and to ensure more precise class results.

   Reaction times vary widely when dilutions extend beyond one volume of Solution A to four volumes of water. Little is gained by assigning solutions more dilute than this.

Part II

5. If one thermometer is used to take the water bath temperature, test-tubes should be in the bath for 5-10 min before students mix the solutions to ensure the "correct" tube temperature. If possible, students can use two thermometers, keeping one in each tube. This will give more reliable results. After the two solutions are mixed, the test-tube containing the mixture should be placed back into the water bath to maintain its temperature.

A temperature of 40 °C should be the highest value assigned, because the HSO_3^- ion concentration is decreased so much that misleading results are obtained. By 50 °C the iodine-starch complex also becomes unstable—another reason to keep temperatures between 0° and 40 °C. Also, the reaction is extremely slow at 0 °C and is difficult to observe and time. Assign each temperature to several groups to obtain some agreement; students will also get an idea of the uncertainty in their measurements from the grouped results.

Teacher-Student Interaction

Move around the laboratory to monitor the uniformity of the solution mixing process. If you have each team do the 5 mL A + 5 mL H_2O reaction first and record the result on the board, you will be able to detect procedural flaws quickly. For *Part II* ascertain the water bath temperature to help prevent overheating. Monitor the central supply of chemical substances to prevent contamination.

Ask students what heating is doing to the molecules. Inquire what molarity the diluted A is as they use it. You can anticipate difficulty with the dilution calculation. Inquire what dilution does to the probability of reactant molecules colliding.

Anticipated Student Results

Part I

Concentration of A	Time (Seconds)	Range of Time (Seconds)
Original (0.01 M)	9	2
4/5 Original (0.008 M)	10	2
3/5 Original (0.006 M)	12	3
2/5 Original (0.004 M)	20	3
1/5 Original (0.002 M)	54	29

Figure 1. Concentration-time graph.

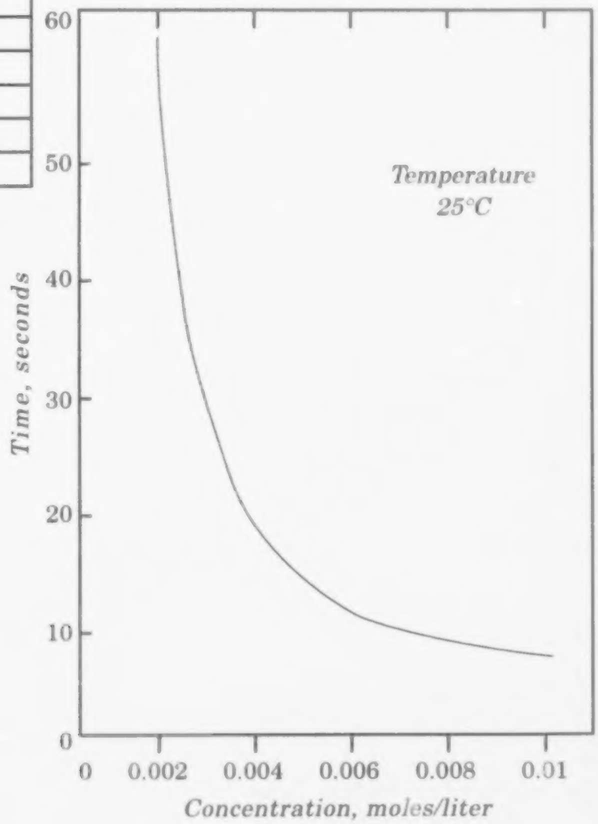


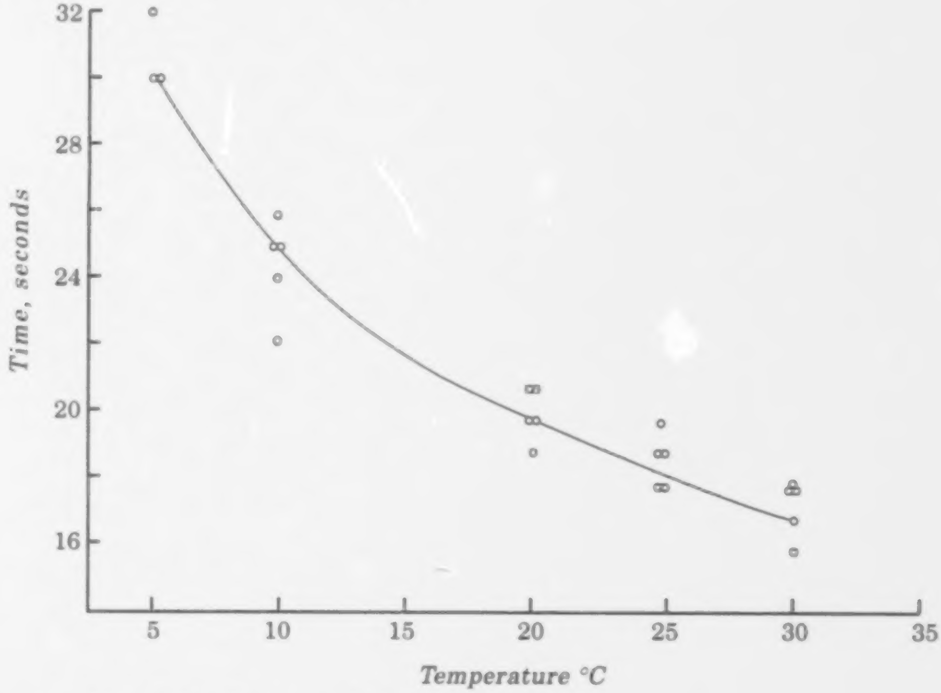
Figure 2. Graph of time vs. concentration (KIO_3).

Part II

Temperature (°C)	Time (Seconds)	Range of Time (Seconds)
5	31	2
10	25	4
20	20	2
25	19	2
30	17	2

Figure 3. Temperature-time data.

Figure 4. Graph of time vs. temperature.



Answers to Data Analysis and Concept Development

Part I

1. $1 \text{ mL sol'n A} \times \frac{1 \text{ L A}}{1000 \text{ mL A}} \times \frac{0.02 \text{ mol A}}{1 \text{ L A}} = 0.00002 \text{ mol A}$

2. $\frac{0.010 \text{ L A}}{0.020 \text{ L sol'n}} \times \frac{0.02 \text{ mol KIO}_3}{1 \text{ L A}} = 0.01 \text{ M KIO}_3$

   $\frac{0.008 \text{ L A}}{0.020 \text{ L sol'n}} \times \frac{0.02 \text{ mol KIO}_3}{1 \text{ L A}} = 0.008 \text{ M KIO}_3$ *etc.*

3. See Figure 2.

Part II

1. See Figure 4.

Answers to Implications and Applications

Part I

1. It is important to keep the total volume at 10.0 mL during the dilution of A for two reasons: to keep calculations simple, and to avoid varying the concentration of Solution B simultaneously.
2. Increasing (decreasing) the concentration decreases (increases) the reaction time. This can easily be seen from the student graph of concentration *vs.* time.
3. A shorter (lower) time means a faster (higher) rate. Therefore rate and time are inversely related.

Part II

1. An increase in temperature decreases the reaction time. You might wish to suggest that better students seek a linear relation between the two variables by plotting various functions of time *vs.* temperature. For the data given, time *vs.* the reciprocal of temperature (1/T) gives a reasonably straight line (where 1/T represents the reciprocal of the *Kelvin* temperature).

   $$\ln k = (-\Delta H/RT) + C$$

2. Since answers may vary considerably, predictions are best reported as ranges. Using data from the graph above, students might answer 35-40 s for 0 °C and 14-16 s for 50 °C.

Post-Laboratory Activities

1. Discuss the results with students. Summarize the laboratory investigation (or have students do it verbally) in terms of relationships between concentration and time (inverse) and concentration and rate (direct), and between temperature and time (inverse) and temperature and rate (direct).
2. Ask students if it is valid to compile results taken over several days for graphing purposes. Discuss problems associated with such a compilation.
3. A computer program such as Vernier's *Graphical Analysis III* can be very helpful to students in their search for a linear relationship involving the variables temperature and time, as discussed in answers to *Implications and Applications*. Functions of the variables can be easily changed within the program with little effort. Consider investigating this as part of class discussion. Results can be displayed on a screen *via* a computer overhead projection device, or on a large-screen television monitor for the entire class to view simultaneously.
4. Depending upon the level of students, you might want to show students the following equation and discuss the reason for delay in appearance of the iodine-starch complex (blue color):

   $$I_2(aq) + HSO_3^-(aq) + H_2O(l) \rightarrow 2I^- + SO_4^{2-}(aq) + 3H^+(aq)$$

   Discuss this reaction in terms of using up iodine molecules, I_2, and preventing them from combining with starch until all HSO_3^- ions are consumed. Only then can I_2 molecules accumulate in sufficient quantity to react with starch molecules and become visible as the blue color.

Extensions

Consider challenging students to find a linear relationship between various functions of the two variables *via* computer program or "by hand." This can provide a challenging exercise for more able students.

Assessing Laboratory Learning

1. Ask students to summarize their findings in a short in-class paper.
2. Provide students a sheet detailing the various steps involved in this reaction (listed in the student laboratory write-up). Ask students to explain why the blue color takes so long to appear. This provides a way to assess their understanding of the *mechanism* of this reaction.

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: The Magnesium-Hydrochloric Acid System—A Study of Reaction Rates

Introduction

Have you ever wondered how chemists slow down reactions that are potentially explosive, or speed up reactions to synthesize a product in a shorter period of time? In this laboratory activity the reaction of magnesium metal with hydrochloric acid,

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

will be used to investigate this problem. The rate may be measured in several different ways. For example, it could be expressed as the volume of H_2 gas produced per second or as the mass of magnesium metal used per second.

Purpose

To design a procedure to measure the rate or speed of the Mg/HCl reaction. You will then identify two factors other than a catalyst to alter the speed of this reaction and will examine each factor quantitatively.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Avoid skin contact with solids and solutions.
3. Hydrochloric acid is a very corrosive reagent. Wash any spills copiously with water. If you need to dilute the hydrochloric acid, remember AAA — "Always Add Acid" to water.
4. No burners should be used during this activity due to the production of potentially explosive hydrogen gas.
5. Dispose of the materials as your teacher directs.

Procedure

1. During the first laboratory period, you will work in a group of four students to investigate the Mg-HCl reaction. Based on these observations, devise a way to measure the rate of this reaction using one of the two approaches described in the introduction. Then propose two different factors that will alter this rate. The group (experimenter) should agree on a procedure in preparation for the second laboratory period. Students within the group should assume these roles: *Chair*, in charge of directing the group and ensuring that all students complete their assigned tasks. *Secretary*, in charge of recording and distributing data to all group members. *Organizer*, in charge of keeping the group on task and maintaining the timetable. *Laboratory Manager*, in charge of obtaining and returning equipment and monitoring safety procedures.
2. Before the second laboratory period, you and your partners individually should:
    a. Sketch a diagram of the apparatus.
    b. List the steps involved in your proposed procedure for investigating effects of the two factors.

c. Create suitable data tables for each factor. You will be expected to complete four different trials for each factor.

Your four-person team will be given:

Magnesium ribbon, 50 cm
3.0 M HCl, 250 mL
Steel wool
Balance
Meter stick
Scissors
Common glassware
Thermometer
Hot plate
Other equipment (if requested in writing the previous day)

3. Your teacher must approve your procedure before you begin your laboratory work.
4. Obtain a piece of magnesium ribbon and clean it with steel wool. Because magnesium has such a low density, it is not practical to weigh small pieces. Therefore, determine the mass of a 50.0-cm length of magnesium ribbon and use this information to calculate the mass of individually measured lengths used in various steps of the activity. For example, if your 50.0-cm length had a mass of 0.40 g, this would be equivalent to 0.40 g/50.0 cm. This equals 0.0080 g/cm or 8.0 mg/cm. (The *mg* unit means milligrams, *not* magnesium.) Record the ribbon lengths and calculated masses.
5. Each group of four students should divide into two pairs. One pair will collect data for one factor, while the second pair collects data for the other factor. At the completion of the laboratory activity, data for both factors should be shared.
6. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. Calculate the reaction rate for each trial performed. Be sure to include units. Convenient units include mg Mg/s, g Mg/s, or mL H_2/s.
2. Prepare a graph of your data for each factor, plotting the independent variable on the *x*-axis and the reaction rate on the *y*-axis.
3. For each graph, predict the reaction rate at another data point not measured in the laboratory activity.
4. What generalizations can you make concerning the effect of each factor you investigated on the reaction rate?
5. How is the rate of a reaction related to the time of the reaction?

Implications and Applications

1. It is often noted that a 10 °C increase in temperature roughly doubles the rate of many chemical reactions. If a reaction takes 20 s at 40 °C, how long would it take at 60 °C?
2. Look up *collision theory* in your text. Use it to explain why each factor tested in this laboratory activity altered the reaction rate.

3. Using a dashed line, sketch what would happen to the appearance of one of the graphs prepared in your data analysis if a catalyst were used in this reaction. Label the graphs.
4. What errors were inherent in the design of your procedure? Explain how each error would affect your measurement of the rate (high, low, or no effect).
5. It has been found that rates are more rapid at the beginning of a reaction than toward the end, assuming the temperature remains constant. Design an experiment using the reaction between magnesium and hydrochloric acid to verify this statement quantitatively.
6. List several instances in everyday life where you face a need to alter the rates of chemical reactions.
7. Did any individuals in your group make unique contributions during this investigation? Explain.
8. Which approach enables you to learn more about a chemical concept—designing your own procedure or following a predetermined set of laboratory directions? Why?
9. Do you prefer developing your own laboratory procedures or following directions that are given to you? Why?
10. List several unanswered questions that arose from this laboratory activity.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: The Magnesium-Hydrochloric Acid System—A Study of Reaction Rates

Major Chemical Concept

Various methods can be used to measure the rate of a chemical reaction. Such procedures allow a quantitative study of the effect of two factors that will alter this rate.

Level

General and honors students. (Basic, if analysis minimized; advanced placement, if used to develop Arrhenius equation.)

Expected Student Background

Students should be able to describe general characteristics of the magnesium/hydrochloric acid reaction, and know the general properties of gases. Students should know how to dilute solutions, calculate solution concentrations (molarities), and prepare and interpret graphs.

Time

Two class periods.

Safety

In addition to routine procedures, warn students of the possibility of cuts from magnesium ribbon, especially while cleaning with steel wool. Use only hot plates to avoid igniting hydrogen. Keep sodium bicarbonate available for acid spills.

Materials (For 24 students working in pairs)

Nonconsumables

12 Balances
12 Hot plates
24 Beakers, 100-mL
12 Meter sticks
Centimeter scale, several
12 Scissors
Copper wire, 15 cm, 22 gauge
12 Thermometers
12 Eudiometer or gas measuring tubes
12 Stoppers, 1-hole, to fit eudiometer tube
12 Graduated cylinders, 25-mL
12 Clocks with second hand, or digital watches with second display

Consumables

Magnesium ribbon, 50 cm
3.0 M HCl, 2000 mL
Steel wool, 12 small lumps

Advance Preparation

1. During the first laboratory period, have small quantities of magnesium ribbon and 3.0 M HCl available at the laboratory stations.
2. To prepare each liter of 3.0 M HCl, slowly add 250 mL 12 M HCl to sufficient distilled water with stirring to produce 1 L solution.

Pre-Laboratory Discussion

Discuss the term *rate* and relate its meaning to the speed of a car. Demonstrate the reaction between magnesium and hydrochloric acid (or a related one) and ask students how they might measure and calculate the rate. Remind students to control variables by changing only one factor at a time. Also remind students that since they will be limited to 50 cm Mg and 250 mL 3 M HCl, they should be judicious in their approach. As part of the assessment process, each student will be asked to evaluate the participation of group members. Pass out the evaluation sheet (see *Student Evaluation of Group Interaction* in *Assessing Laboratory Learning*) so students can see which behaviors will be noted during the laboratory activity. Students may be allowed to select only one of the behaviors listed, in keeping with cooperative learning procedures, if you so choose.

Teacher-Student Interaction

Remind students not to heat the hydrochloric acid above 45 °C and to watch for splattering. While students are involved in experimentation, watch for proper laboratory techniques and safety concerns; also check for proper recording of data and use of significant figures (see check list in *Assessing Student Learning*). Also monitor that their approved procedure is being followed.

Anticipated Student Results

Be ready for a variety of experimental designs, some of which are excellent and others which will require modification and teacher guidance.

The most common designs are:

1. Cut the magnesium into various lengths and time how long they take to dissolve. (This is probably the simplest and best but presents the problem of keeping the magnesium submerged.)
2. Place magnesium in a eudiometer or inverted graduated cylinder and time the gas evolution at intervals. (This has the problem of getting the magnesium into a uniform acid concentration without gas loss.)
3. Observe the temperature increase of the solution with time. (Since ΔT is small, a very sensitive thermometer is necessary for success.)
4. Observe the change in pH as acid is consumed. (Unless volume of acid is small, this is a small effect. It also requires a pH meter.)

Factor 1: Effect of Decreasing the HCl Concentration on the Rate

Typical data obtained using a 5.0-cm length of Mg ribbon, 20.0 mL 3M HCl in a small beaker, measuring the total time for all magnesium to be consumed:

Concentration HCl (M)	3.0	2.0	1.5	0.8
Time (s)	8	11	18	36

Typical data obtained using a 5.0-cm length of Mg ribbon encased in a copper wire cage, 15 mL HCl added to a eudiometer tube and filling the remainder of the tube with water, measuring the total time to generate 16.0 mL hydrogen gas:

Concentration HCl (M)	3.0	2.0	1.5	0.8
Time (s)	127	210	230	355

Factor 2: Effect of Increasing the Temperature on the Rate

Typical data using a 5.0-cm length of Mg ribbon, 1.0 M HCl in a small beaker, measuring the total time for all magnesium to be consumed.

Temperature (°C)	5.0	15.4	23.5	32.8	43.5
Time (s)	168	107	66	45	32

Factor 3: Effect of Increasing the Surface Area on the Rate

Typical data using 10.0-cm length of Mg ribbon, 1.0 M HCl in a small beaker, measuring the total time for all magnesium to be consumed when the strip is rolled into a tightly coiled piece, a medium-size coil and a loosely coiled piece.

Circumference of Mg coil (cm)	2.0	2.9	4.8
Time (s)	148	111	95

Answers to Data Analysis

1. Rate can be determined by dividing volume of hydrogen gas produced by time involved or mass of Mg used by time involved. For example, using the data above:

   5.0 cm Mg used x (0.41 g Mg/50.0 cm Mg)/8 s = **0.005 g Mg/s**

   or 16.5 mL H_2 produced/127 s = **0.130 mL H_2/s**

2. Graphs:

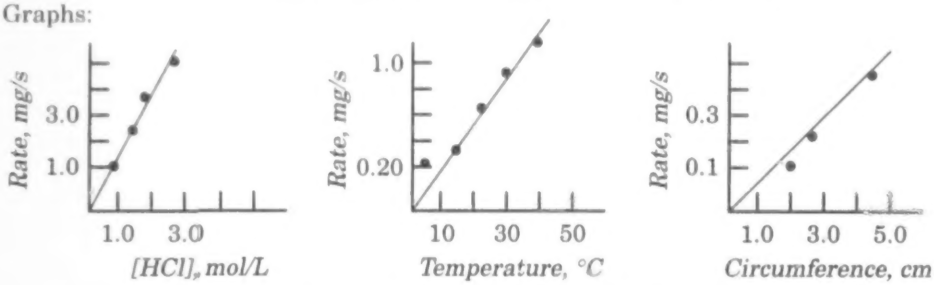


Figure 5. Graphs of rate versus (a) concentration of HCl, (b) temperature, (c) circumference.

   A computer program for curve-fitting such data could be used in this analysis. One example: *Graphical Analysis III* (Vernier Software).

3. Students use their graphs to estimate two other data points.
4. As the concentration of a reactant increases, the reaction rate increases. As the temperature of the reaction increases, the rate increases. As the surface area increases, the rate increases.
5. The rate of a reaction is inversely proportional to the time of the reaction.

Answers to Implications and Applications

1. 5 s
2. As the concentration of reactants decreases, the number of effective collisions per unit time decreases; hence the reaction rate decreases. As the temperature increases, more molecules possess sufficient energy of activation that causes the reaction rate to increase. As the surface area increases, the total number of effective collisions per unit time increases; hence the reaction rate increases.

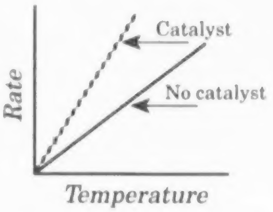


Figure 6. Graph of rate vs. temperature.

3. Using a catalyst, the graph would resemble this example.
4. Possible errors:
    a. Since the reaction is exothermic, the temperature (and hence the rate) would increase from this factor alone.
    b. The temperature is difficult to control and varies when using the eudiometer tube.
    c. Other sources of error such as nonuniform strips of Mg, spillage, improper dilution, *etc.*
5. Sample answer: Measure the volume of hydrogen gas produced from the reaction in a eudiometer tube over several minutes, taking volume readings at 10-s intervals. Record both the volume and the time. The volume produced initially is greater than the volume produced later in the reaction for a given time interval.
6. a. Using a platinum catalyst in contact lens sterilization vials to decompose hydrogen peroxide solutions.
    b. Using kindling in campfires.
    c. Preserving perishable foods in a freezer or refrigerator.

7-10. Use teacher discretion.

Post-Laboratory Activities

1. Group students according to the factor they chose to investigate; have them share their data and reach collective conclusions to be shared with the entire class.
2. Discuss various unique experimental designs and features offered by each. Discuss the importance of reaction rates in real-world applications.

Extensions

This laboratory activity provides an excellent lead in to rate law expressions.

Assessing Laboratory Learning

Quiz Items

1. What is meant by the term "rate of reaction"?
2. What experimental data must be collected to determine rate of a reaction?
3. You have three samples of solid calcium carbonate, $CaCO_3$, of equal mass, cut in the following pieces: (a) one perfect cube, 2.0 cm on a side; (b) eight perfect cubes, each 1.0 cm on a side; (c) 64 perfect cubes, each 0.50 cm on a side. Assume you are interested in determining the effect of total surface area on the rate of this reaction:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g).$$

You decide to use the apparatus in Figure 7.

a. What data would you need to collect?

b. Explain how you would use these data to calculate the rate of this reaction.

c. Predict how the rate would be affected as the surface area of $CaCO_3$ is increased in this reaction.

d. Explain your prediction in Part 3c using collision theory.

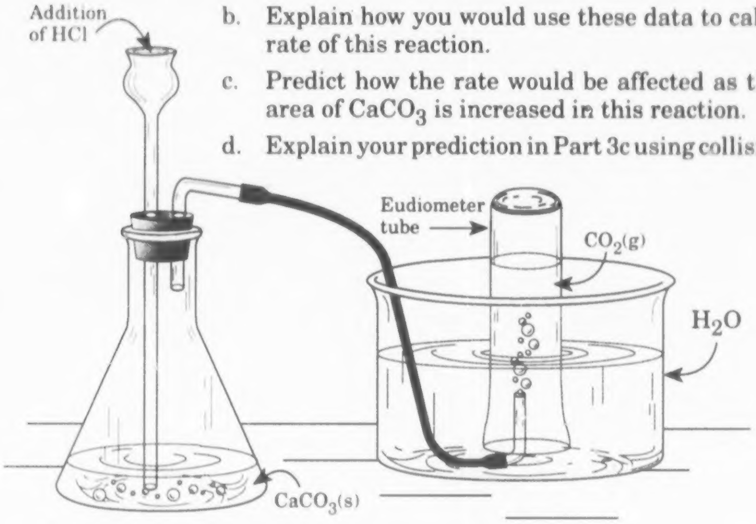


Figure 7. Apparatus for collection of $CO_2(g)$.

Laboratory Report

The formal report should include a brief purpose (in student language), description of procedure, data table with all pertinent values labeled, answers to Data Analysis/ Implications and Applications, and conclusion/summary of concepts.

Pencil/Paper Items

1. *[This assessment item could be administered as a teacher demonstration, possibly videotaped for ease of repetition.]* Hydrogen peroxide decomposes at a reasonable rate in the presence of a catalyst according to this equation:

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

The catalyst is placed into a tissue-paper bag (see transparency). When it drops into the solution of hydrogen peroxide in the flask, the released oxygen gas is collected in the eudiometer tube. If you were interested in studying the effect of changes in H_2O_2 concentration on the reaction rate,

   a. What data would be convenient to collect to determine the rate of this reaction?

   b. How would you calculate the rate of this reaction from data described in 1a?

   c. What variables must be controlled in order to complete this investigation?

2. From data obtained in the above activity, a student found that the rate of reaction increased as the concentration of hydrogen peroxide increased.

   a. Sketch the general shape of the graph that might be obtained from these data, with the rate of reaction on the y-axis and the hydrogen peroxide concentration on the x-axis. Label the axes.

   b. Redraw the graph from 2a and with a dotted line, and indicate the graph that would be expected if the reaction were completed without a catalyst. Label the two plotted lines.

   c. Sketch the graph that might be obtained by plotting the time of reaction on the y-axis and the H_2O_2 concentration on the x-axis. Label the axes.

   d. Explain the effect of increasing the H_2O_2 concentration on the reaction rate in terms of collision theory.

3. Rates are observed to be more rapid at the beginning of a reaction than toward the end, assuming the temperature is held constant.

   a. Design an experiment using the reaction between magnesium and hydrochloric acid to gather data to verify this observation.

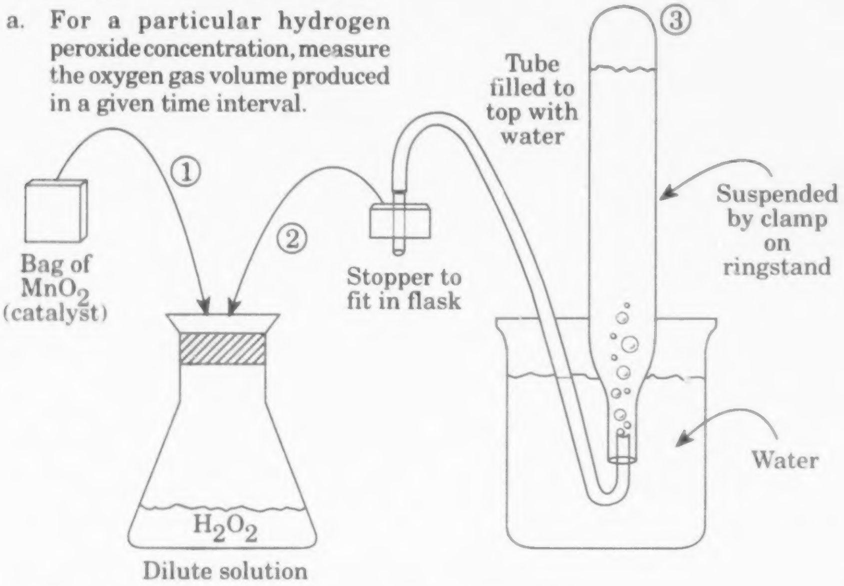
   b. Indicate what data need to be collected in this experiment.

   c. Explain this observation in terms of collision theory.

4. Describe three ways the rate of a reaction can be increased.

Answers to Paper/Pencil Assessment

1. a. For a particular hydrogen peroxide concentration, measure the oxygen gas volume produced in a given time interval.


③
Tube filled to top with water
①
②
Bag of MnO_2 (catalyst)
Stopper to fit in flask
Suspended by clamp on ringstand
Water
H_2O_2
Dilute solution

Figrue 8. Decomposition of hydrogen peroxide.

   b. Divide the volume of oxygen gas produced by the time interval.

   c. Temperature, air pressure, volume of hydrogen peroxide used.

2. See a, b, and c.

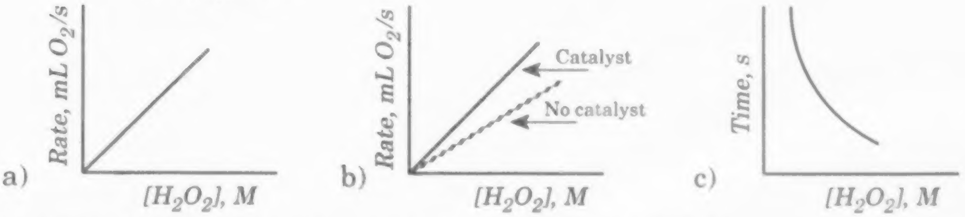


Figure 9. Graphs of reaction rate vs. concentration of H_2O_2.

3. a. See answer to Question 6 in *Implications and Applications* section.
   b. Data: volume of hydrogen gas produced over constant time periods (such as 10-s intervals). Record both volume and time.
   c. There are more collisions among reactant molecules per unit time at the beginning of a reaction, due to their higher concentration. Toward the end of a reaction, fewer reactant molecules remain—there are fewer collisions and the rate decreases.
4. Increase temperature, concentration of reactants, surface area; add a catalyst.

Check List

Teacher Evaluation

1. Procedure written out prior to class
2. Data sheet prepared prior to class
3. Use of proper laboratory technique, recording of data (attending to proper use of significant digits)
4. Use of proper safety precautions (goggles, dilution procedures, heating, *etc.*)
5. Clean up

Student Evaluation of Group Interaction

(To be completed anonymously by each member of the group)

1. Listening to others
2. Communicating ideas
3. Contributing to overall design of experiment
4. Helping others where possible
5. Overall participation

Possible scoring scheme: 3 points: Excellent
2 points: Satisfactory
1 point: Poor

DEMONSTRATIONS

Demonstration 1: Effect of Surface Area on Rates of Reaction

Materials

2 Steel wool pads, size 0000
Lycopodium powder, 5 g
Iron powder, 20 g
Iron filings, 20 g
Iron rod
1 Sugar cube (generally 3-4 g)
Confectioner's or powdered sugar, 3-4 g (must match weight of cube)
Granulated sugar, 3-4 g (must match weight of cube)
Water, 600 mL
Flint striker
2 Plastic squeeze bottles, small, with outlet tube to bottom
4 Beakers, 400-mL
Rubber tubing with small piece of fire-polished glass at each end
2 Burners
Safety glasses
Powdered zinc, 2-3 g
Mossy zinc, 3-4 small pieces
6 M Hydrochloric acid, HCl, 50 mL

Directions

Some of these procedures are most effective when performed in a darkened room.

1. Fill one of the two small plastic squeeze bottles about 2 cm deep with iron powder; fill the other bottle 2 cm deep with lycopodium powder. Light the two burners and place them about 10 cm apart, lined up to face an area *away* from students. Spray the solids one at a time into the burner flames. Practice this procedure before doing the demonstration.
2. Place a small pile of lycopodium powder on a clay plate or insulated pad and direct a burner flame onto the pile. Repeat with iron filings.
3. Place a pad of 0000 steel wool into a 400 mL beaker. Fluff it up to form a depression in the middle. Hold the flint striker over the steel wool and ignite the steel wool. Blow into the glowing metal at first with your mouth to get the reaction started, and then continue blowing with the use of the rubber tubing with glass tubing at each end.
4. Contrast the ignition of the iron filings (these will sparkle in the flame) and the iron pad to the burning of iron powder as seen in Part 1.
5. Contrast the dissolving in water of equal masses of sucrose in the form of confectioner's, granular, and cubes, by placing approximately equal amounts of each of three varieties in each of three 400-mL beakers containing approximately 200 mL water.
6. Place a shallow layer of 6 M HCl in each of two beakers on an overhead projector. To the first beaker, add a small chunk of mossy zinc and observe the slow evolution of bubbles. Sprinkle powdered zinc into the second container and compare this vigorous bubbling to the slow bubbling in the first. Be prepared to quench the reactions by pouring in excess water at the close of the demonstration.

Discussion

1. Lycopodium powder is composed of yellowish powdery spores of certain club mosses used in fireworks, explosives, and as a covering for pills. Lycopodium means "wolf foot"—the plant roots are claw-shaped. In Greek, *lykos* means wolf and *pod-* is the stem for foot. It is a mixture that includes organic compounds.

Safety

1. Try all parts of this demonstration in advance to establish the safest reaction conditions.
2. Some people may be allergic to lycopodium powder. Be sure to provide adequate ventilation.
3. Wear protective goggles for all chemical demonstrations.

Demonstration 2. Catalysis and Activated Complex Formation

Materials

0.3 M Potassium sodium tartrate (Rochelle's salt) solution, 600 mL (50.8 g $KNaC_4H_4O_6 \cdot 4H_2O$ per 600 mL solution)

6% Hydrogen peroxide solution, 200 mL (Use Clairoxide™ hair developer (6%), available from drug stores, supermarkets, or beauty supply companies.)

0.3 M cobalt(II) chloride, 30 mL (2 g $CoCl_2 \cdot 6H_2O$ per 30 mL solution)
3 Beakers, 600-mL
2 Beakers, 30 mL (50-mL beakers will suffice)
2 Watch glasses large enough to cover 600-mL beakers
3 Thermometers
Hot plate, large

Directions

1. To each 600 mL beaker add 200 mL potassium sodium tartrate solution and 65 mL 6% hydrogen peroxide solution.
2. Cover each beaker with a watch glass and heat on the hot plate until the temperature is between 65 °C and 70 °C. *Do not exceed this temperature.* While the large beakers are heating, divide the 30 mL cobalt chloride solution equally between the two smaller beakers.
3. Add one 15 mL portion of cobalt chloride solution to one of the larger beakers, cover and observe. Keep the second sample of cobalt chloride solution plainly visible as a color comparison; have students note the color of the original cobalt chloride solution (pink).

Discussion

Catalysis is demonstrated by the vigorous reaction in the beaker containing the $CoCl_2$ catalyst. The foaming is due to production of $CO_2(g)$ bubbles. The other, uncatalyzed beakers show little or no visible bubbling.

Activated complex formation is demonstrated by the color change from pink to green; the green color indicates production of a Co(III) complex, although that is not the activated complex. (Activated complexes typically exist for too brief a time to be observable.)

When the reaction is complete (no more bubbling), the catalyst is regenerated. This is shown by formation once again of pink color, indicating the regeneration of the (pink) $CoCl_2$ catalyst.

To emphasize that the catalyst has been regenerated, add a small amount of spent solution to the second 600-mL beaker, cover with a watch glass, and observe. You will note a repeat of fairly rapid bubbling, as well as formation of new green product in this second beaker. The third beaker serves as a control to demonstrate the slowness of the uncatalyzed reaction.

Demonstration 3. Catalysis

Materials

Powdered zinc, Zn, 1g
Iodine crystals, I_2, 4g
Crucible or evaporating dish
Squeeze bottle of water

Directions

Mix the powdered zinc and iodine crystals together in a fume hood and wait for a reaction to begin. The reaction will not occur, no matter how long you wait. Now add just a few drops of water to the mixture and stand back. Immediately a large puff of purple iodine vapor rises, indicating that considerable heat has been generated by the reaction.

Discussion

Why didn't the reaction begin without the water? Apparently the activation energy is too large. Adding the water provides a new path unavailable before.

(1) $Zn(s) + I_2(s) \rightarrow ZnI_2(s) \qquad \Delta H = -208 \text{ kJ/mol}$

(2) $ZnI_2(s) + H_2O(l) \rightarrow ZnI_2(aq) \qquad \Delta H = -56.23 \text{ kJ/mol}$

When the reaction begins in the presence of water, sufficient thermal energy is produced to melt and vaporize the iodine. The vaporization of iodine accounts for the violet smoke. Since water can be recovered after the reaction, it acts in this reaction as a catlayst. Reactions between solids at room temperature are frequently slow since the reaction can only occur at the solid surfaces. The rate increases rapidly with time due to the temperature increase that follows the exothermic reactions, at which time the water may no longer be necessary to maintain a high reaction rate. The initial thin coating of $ZnI_2(s)$ on the zinc is solubilized by the drop of water.

Safety Considerations

Be wary. This reaction is extremely rapid. The reaction should be done in a fume hood, or at least with very good ventilation, as iodine fumes can be choking to students (and teacher).

Demonstration 4: Light Sticks as Rate Indicators

Materials

2 Beakers, 400-mL
2 or 3 Light sticks (Cyalume™ or others; two minimum)
Source of hot water
Water bath, ice cold

Directions

Open two light stick packages and initiate the reaction by bending the tubes until the inner glass ampule breaks, thus mixing the reactants. Allow students to observe that both sticks emit the same intensity of light. Now place one stick in a hot water bath (not hotter than 80 °C) and one in the ice-water bath. (You may decide to keep a third light stick at room temperature as a control.) Allow several minutes for the two light sticks to equilibrate temperatures with the baths. Have students observe the difference in intensity (brightness). The light stick in hot water will glow very brightly while the tube in cold water will have dimmed considerably.

Discussion

Discuss temperature's effect on reaction rate. It is also good to follow up this demonstration with the question: "Which tube will emit light longer, and why?" Discuss total amounts of reacting substances in the tubes and rate at which those materials are consumed. This also leads to the notion of keeping the sticks in the freezer overnight to preserve the light for another day, which most students already understand.

Safety

The light sticks contain nontoxic substances, according to the manufacturers. Nevertheless, it is wise to discard the used light sticks without opening them.

Demonstration 5: The "Stoplight" Reaction

Materials

1 M Sodium hydroxide, NaOH (4 g NaOH per 100 mL solution)
Dextrose solution (2.5 g dextrose per 100 mL solution)
1% Indigo carmine solution (0.1 g per 10 mL solution)
Erlenmeyer flask, 250-mL or 500-mL
Stopper to fit flask

Directions

Prepare the sodium hydroxide and dextrose solutions separately. Just before starting the demonstration, mix equal volumes (*e.g.*, 50 mL) of the two solutions in the Erlenmeyer flask. Add a few drops of indigo carmine solution (just enough to give the solution a distinct color). Stopper the flask and let it sit undisturbed for several minutes. The color should now be yellow-orange. First, swirl the flask gently; it should change to a red-orange color. Now give the flask one or two quick, hard shakes. The solution will turn yellow-green. Upon standing, the flask will change back to the red-orange color, and then back to the original yellow-orange color. Repeated shakings will duplicate this color change sequence.

Discussion

By varying the number and severity of shakes, and by noting relative rates of the color changes, students may begin to understand the factors affecting this reaction. Consider asking students to suggest such variations in the laboratory investigation (and others) to evaluate their understanding of factors that affect reaction rates.

Safety

NaOH solution is rather concentrated (1 M) and caustic. Use care when handling the solution. Discard by rinsing down the drain, flushing with plenty of water. The dextrose solution can be disposed of in a similar manner. The indigo carmine solution will stain clothes; use care. The mixed solution will decompose with time; it is best to prepare a fresh mixture near the time for the demonstration.

Reference

Ferguson, H. W., Schmuckler, J. S., Caru, A. N., and Siegelman, I. J. (1970). *Laboratory Investigations in Chemistry* (Teacher's Edition). Silver-Burdett.

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. What is a reaction rate? *[A change of concentration of a reactant or product per unit time.]*
2. What causes a rate to be slow or fast? *[Nature of reactants (solid, liquid, gas, solution); number of bonds that must be broken and reformed; number of bodies that must collide for reaction to occur.]*
3. What factors affect the rate of a chemical reaction? *[Concentration, temperature, pressure (gaseous system), surface area (heterogeneous system), catalyst.]*
4. What is a reaction mechanism? *[A series of reaction steps on a molecular level whose sum is equal to the balanced overall equation and which are in agreement with the experimentally determined rate equation. Example:*

$$\begin{array}{llcl}
\textit{Step 1} & HCOOH + H_2SO_4 & \rightarrow & HCOOH_2^+ + HSO_4^- \\
\textit{Step 2} & HCOOH_2^+ & \rightarrow & HCO^+ + H_2O \\
\textit{Step 3} & HCO^+ + HSO_4^- & \rightarrow & H_2SO_4 + CO \\
\hline
\textit{Overall} & HCOOH & \rightarrow & H_2O + CO
\end{array}$$

Without the concentrated H_2SO_4 the reaction is very slow:

$$HCOOH \rightarrow H_2O + CO]$$

5. Define each term, giving an example of each:

   a. Catalyst *[A species that speeds up the rate of a reaction by providing a new mechanism with a lower activation energy. Example, H_2SO_4 in Question 4 is a catalyst.]*

   b. Reaction intermediate *[A product of one reaction that is consumed in a following reaction in a process that proceeds through several steps. Examples, $HCOOH_2^+$ and HCO^+ in the mechanism in Question 4.]*

   c. Activated complex *[A specific arrangement of atoms that has the highest energy during the rearrangement of the atoms of the reactants to those of the products. It is also called the transition state.]*

   d. Inhibitor *[A species that bonds with reacting species, thus lowering the overall rate of a reaction. Example, calcium propionate retarding the spoilage of bread.]*

6. Using the kinetic molecular (collision) theory, explain how the rate of a reaction is affected by:

   a. increasing the concentration of a reactant. *[Increasing the concentration provides a greater number of reacting species, resulting in greater number of collisions per unit time, and therefore produces a faster reaction rate.]*

   b. adding a catalyst. *[A catalyst provides a new mechanism for the reaction that has a lower activation energy barrier, yielding a faster rate.]*

   c. raising the temperature. *[Higher temperature produces a faster average movement of reacting species. Faster particles collide more frequently and with greater impact and result in more effective collisions, producing a faster reaction rate.]*

   d. using a spray of liquid or a powdered solid. *[Smaller solid particles or liquid drops provide greater reactive surface areas, giving rise to more collisions per unit time and a faster reaction rate.]*

   e. stirring the reactant mixture. *[More thorough mixing of reactants enhances fruitful collisions and reduces reaction time, resulting in faster rate.]*

7. Why do chemists assume most reactions occurring at measurable rates involve only bimolecular collisions? *[The probability of anything more than a two-body collision is extremely rare and thus would contribute negligibly to the overall reaction rate. A termolecular collision involves three molecules arriving at the same x, y, z coordinates at exactly the same time. A termolecular collision would be more probable in an enzymatic active site reaction.]*

8. The balanced equation representing a chemical reaction rarely indicates the mechanism of that reaction. Why? *[Equations merely show stoichiometric relationships between reacting species and products formed; they do not represent the mechanism of the reaction and do not indicate or even hint at the number of bodies involved in the actual overall reaction.]*

9. Why is experimentation the only way to learn about the rate of a reaction not previously studied? *[As previously mentioned, an overall equation cannot suggest a mechanism for most reactions. Experiments are needed to determine the nature of the reaction intermediates and the reaction order of each reactant.]*

Counterintuitive Examples and Discrepant Events

1. The equation $4HBr + O_2 \rightarrow 2H_2O + 2Br_2$ seems like it would represent an extremely slow reaction, if it could happen at all, due to the extremely low probability of a single simultaneous five-body collision. Yet the reaction happens at a fairly moderate rate.

Explanation

The overall balanced equation indicates only the quantities of each substance in the initial and final states of the chemical system. It does not usually indicate the actual steps (reaction mechanism) that occur on a molecular level, but does indicate the sum of these individual steps. Rate equations that describe how specific reactions depend on the concentrations of reactants must be determined experimentally, not by considering the overall balanced equation. A reaction mechanism for the reaction of HBr with O_2 helps to explain how this reaction can happen by a series of two-body collisions. The accepted mechanism for the reaction is:

Step 1	HBr (1)	+	O_2	→	~~HOOBr~~		
Step 2	~~HOOBr~~	+	HBr (2)	→	~~2 HOBr~~		
Step 3	~~HOBr~~	+	HBr (3)	→	H_2O	+	Br_2
Step 4	~~HOBr~~	+	HBr (4)	→	H_2O	+	Br_2
Overall	4HBr	+	O_2	→	$2H_2O$	+	$2Br_2$

As you can see, the actual reaction does occur by a series of two-body collisions, a much more probable scenario than a single, simultaneous five-body collision.

2. Another example of this type of counterintuitive phenomenon is the following equation:

$$2H^+ + 2Br^- + H_2O_2 \rightarrow Br_2 + 2H_2O$$

Explanation

The proposed mechanism for this reaction is:

Step 1	H^+	+	H_2O_2	$\rightleftharpoons$	$H_3O_2^+$			(Bimolecular)
Step 2	Br^-	+	$H_3O_2^+$	→	H_2O	+	HBrO	(Bimolecular)
Step 3	Br^-	+	HBrO	→	Br_2	+	OH^-	(Bimolecular)
Step 4	H^+	+	OH^-	→	H_2O			(Bimolecular)

Overall $2H^+ + 2Br^- + H_2O_2 \rightarrow Br_2 + 2H_2O$

Once again, the mechanism involves only two-body collisions.

3. A container filled with a 2:1 mole ratio of hydrogen gas and oxygen gas has been stored in a safe place for 20 years. It is a known fact that the reaction of two moles of hydrogen gas with one mole of oxygen gas is highly exothermic, releasing 571 kJ of heat. Explain why the two gases in the stored container have not reacted even after 20 years.

Explanation

Many chemical reactions are very slow because the activation energy barrier, the minimum energy needed to initiate the reaction, is very high. This usually means that one or more chemical bonds in reactants (that must be broken to initiate the reaction) has a high bond energy. In this case, bond dissociation energies are: H-H, 435 kJ/mol and 0-0, 494 kJ/mol. The energy of attraction within reactant molecules in this reaction is very high as a consequence of the strong bonds within H_2 and O_2. This makes these two molecules very stable and difficult to split apart. Finely divided platinum or a platinum gauze lowers the energy of activation, allowing the two species to react more easily, thus increasing the reaction rate. Here is the proposed mechanism:

Step 1 $Pt(s) + O_2(g) \rightarrow PtO_2(s)$
Step 2 $PtO_2(s) \rightarrow Pt(s) + 2O\cdot$
Step 3 $O\cdot + H_2(g) \rightarrow H_2O(g)$
Step 4 $O\cdot + H_2(g) \rightarrow H_2O(g)$

Overall $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

(NOTE: O· is a free radical, Pt is a catalyst, PtO_2 *is a reaction intermediate.)*

Metaphors and Analogies

1.

Topic	Quantity	Time	Rate
Respiration	Breaths	Minute	60 breaths/min
Salary	Dollars	Week	$400/week
Heartbeat	Beats	Minute	75 beats/min
Speed of light	Miles	Second	186,000 miles/s
Car speed	Miles	Hour	55 miles/h
Keyboard speed	Words	Minute	60 words/min
Shower flow	Gallons	Minute	3 gal/min
Speed of sound	Feet	Second	1100 ft/s

Figure 10. Rates in everyday life.

2. **Mechanical Analogy of Rate Terminology with Funnels**

Equipment

1 L food dye/water solution; at least four funnels with a wide range of bore diameters (the stems may be constricted by fire polishing or drawing the glass stems to a finer opening); six iron rings to hold the funnels; a large-size "Y" connector tube; two 250-mL graduated cylinders; a timer (seconds); a beaker to catch liquid outflow; one tall ringstand (see Figure 11).

Procedures

1. Using each funnel separately, measure the time (in seconds) to collect 250 mL , or measure the volume of solution (in milliliters) collected in a prescribed time (*e.g.*, 15 seconds).
2. Set up the funnels in series and repeat the options in Step 1 for liquid passing through all four funnels.

3. To illustrate concurrent (simultaneous) reactions, pour the colored liquid into a funnel connected to an inverted "Y" tube that, in turn, feeds two other funnels of differing bore diameters situated above two 250-mL graduated cylinders. The comparative rates passing through the two funnels of different bores can be observed or measured quantitatively.
4. If the system in Step 3 is allowed to come to equilibrium, a contrast can be made between kinetics and equilibrium in a chemical system.

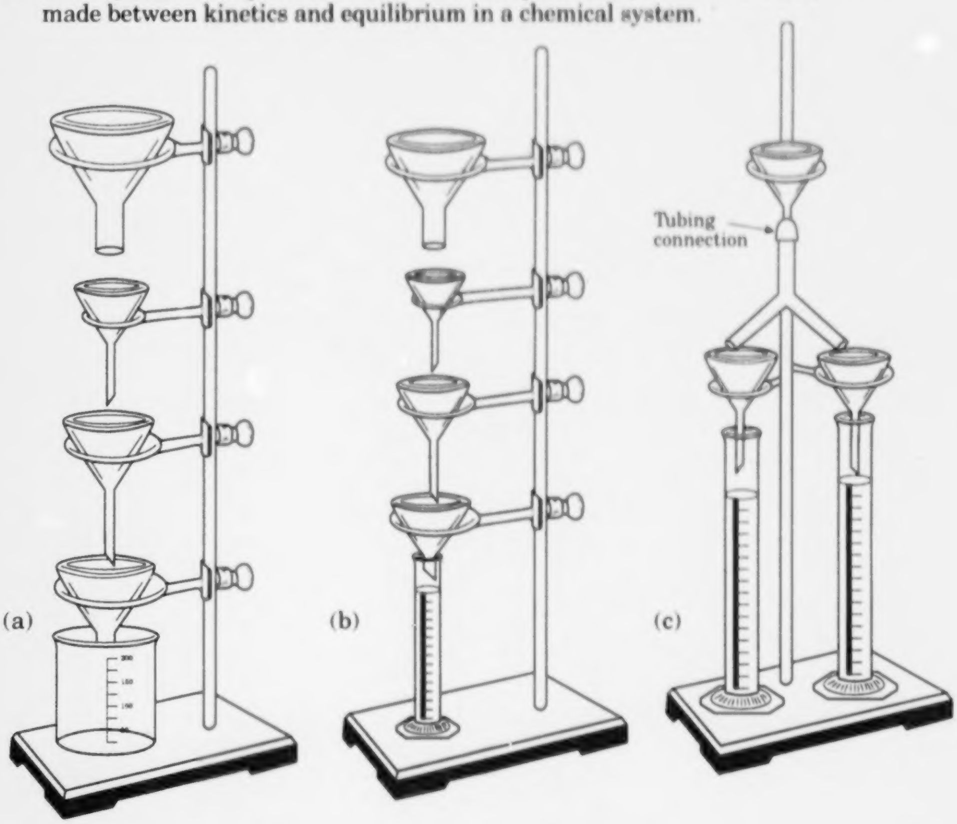


Figure 11. Mechanical analogy for rate terminology: (a) consecutive reactions (qualitative), (b) consecutive reactions (quantitative), (c) concurrent (simultaneous).

Results

Typical data for the flow rate through single funnels and all four in series are provided in Figures 12 and 13.

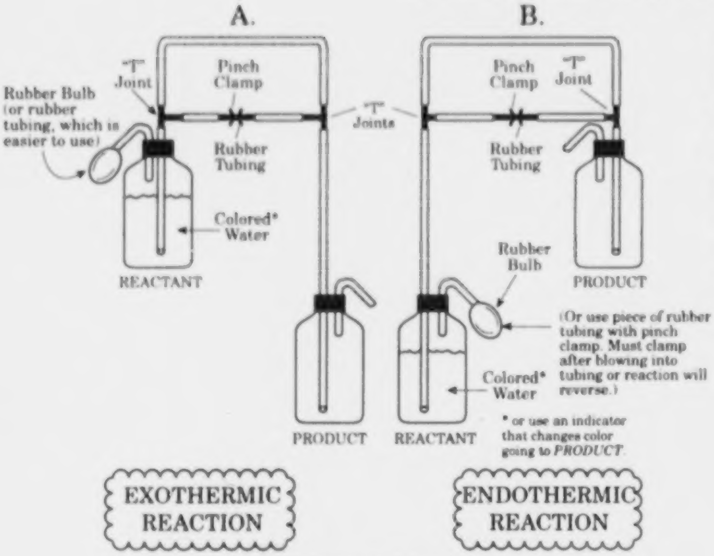
Funnel No.	Volume (mL) Collected in 15 s	Rate (mL/s)
1	420	28
2	350	23
3	55	3.7*
4	470	31
All four funnels	58	3.9
*Rate-determining step		

Figure 12. Volume (mL)

Funnel No.	Trial 1	Trial 2	Average	Rate (mL/s)
1	9.2	8.9	9.1	27.0
2	11.2	11.3	11.2	22.0
3	65.8	62.6	64.2	3.9*
4+	6.0	6.2	6.1	41.0
All four funnels	61.4	62.9	62.1	4.0
+Not the same as Funnel 4 in Part 1 *Rate-determining step				

Figure 13. Times (s) to collect 250 mL.

3. "The Glass Menagerie": A Mechanical Analogy for Exothermic and Endothermic Reactions, Activation Energy, Catalyzed *vs.* Uncatalyzed Reactions, Reaction Reversibility, and Energy Barriers to Reaction.


A.
B.
"T" Joint
Pinch Clamp
Rubber Bulb (or rubber tubing, which is easier to use)
Rubber Tubing
Colored* Water
REACTANT
"T" Joints
PRODUCT
Pinch Clamp
"T" Joint
Rubber Tubing
Rubber Bulb
PRODUCT
(Or use piece of rubber tubing with pinch clamp. Must clamp after blowing into tubing or reaction will reverse.)
Colored* Water
REACTANT
* or use an indicator that changes color going to *PRODUCT*.
EXOTHERMIC REACTION
ENDOTHERMIC REACTION

PRINCIPLES ILLUSTRATED

1. Exothermic Reaction, ΔH
2. Endothermic Reaction, ΔH
3. Activation Energy
4. Energy Barrier
5. Catalyzed Reaction
6. Uncatalyzed Reaction
7. Reaction Reversibility

Figure 14. Reaction analogy apparatus.

Equipment

Two large transparent or translucent bottles (*i.e.*, 6-lb acid bottles); two 2-hole rubber stoppers; glass tubing as shown in Figure 14; two glass or metal "T" joints; pinch clamp; siphon pump; rubber bulb; pieces of rubber tubing; acid-base indicator with a little sulfuric acid (both on reactant side); and a few flakes or pellets of sodium or potassium hydroxide (on product side).

Procedure

1. Set up apparatus as indicated in Figure 14. (*NOTE: Setup B is just Setup A reversed in direction.*) Select dimensions to facilitate easy storage and reuse. Rest the higher bottle on a box. Add the dilute sulfuric acid and indicator to the "reactant" bottle. Clamp the lower tube (catalytic route) shut (or use rubber bulb). Initiate the "reaction" by blowing into the rubber tube until the liquids flow into the "product" bottle. Discuss the appropriate concepts. Now open the clamp to allow the reaction to follow the catalytic route with lower activation energy (you could use a wider-bore piece of tubing for the catalytic route, which would transfer liquid faster). Discuss inhibitors as well as catalysts.

2. When the product bottle becomes filled, blow into the rubber tubing (or use rubber bulb). You will need to clamp it off *if* you want to prevent the "reaction" from reversing. Endothermic reactions are nonspontaneous and require continued input of energy to continue. (Actually, in chemical changes the change in free energy (that is, $\Delta G = \Delta H - T\Delta S$) is the criterion to start the endothermic reaction.) Discuss a few examples of endothermic reactions. Next open the clamp, allowing the reverse catalytic reaction to occur. Remove the siphon pump and note that the reaction changes direction when either catalyzed or uncatalyzed. *NOTE: It may simplify things to use two three-hole stoppers and separate longer tubing for the catalyzed route. You would need a bottle that would accommodate the larger stopper.*

NOTE: In place of rubber bulbs consider using pieces of rubber tubing with pinch clamps. You only have to blow into the tubing of the upper bottle to start the exothermic reaction. You would need to blow continuously into the tubing attached to the lower bottle to keep the "endothermic reaction" going or use a pinch clamp.

4. **Marble Shoot.** Have students in laboratory pairs stand around the laboratory bench and roll a marble from two different directions, at a central target. Discuss the likelihood that the two marbles will collide. Then have three students each roll a marble toward the same target and observe the types of collisions that result. Invariably students will observe a series of two-body collisions, rather than a single simultaneous three-body collision, just as with real molecules in reaction mechanisms. If you can control the group, you might next try four marbles, but watch out!
5. Effective *vs.* ineffective collisions can be illustrated by the analogy of a head-on collision between two autos, damaging both (the "effective" collision) *vs.* a gentle tap between bumpers of two autos ("ineffective" collision).

TIPS FOR THE TEACHER

Language of Chemistry

activated complex specific arrangement of atoms that has the highest energy during the rearrangement of the atoms of the reactants to those of the products. It is also called the transition state.

activation energy minimum quantity of energy needed for reacting species to initiate a chemical reaction (or to form the activated complex).

active sites locations at which catalysis occurs.

Arrhenius equation equation that relates the rate constant of a reaction to temperature and activation energy.

bimolecular reaction elementary process involving the collision of two species.

catalyst species that increases the rate of a reaction without being used up in the reaction. The catalyst may or may not be used along with the reactants, but if it is consumed, it is always regenerated at the end of the reaction. The catalyst offers a reaction mechanism with a lower activation energy.

chemical kinetics study of speeds or rates at which chemical reactions occur.

collision theory theory states that, in order for a reaction to occur, reactants must collide with sufficient force and proper geometry.

elementary steps series of simple reactions that represents the progress of the overall reaction at the molecular level.

enzyme biological catalyst.

first-order reaction reaction whose rate depends on reactant concentration raised to the first power.

half-life time required for one-half of a reactant to be consumed in a chemical reaction. In a nuclear decay process, it represents the time required for one-half the atoms in a given sample of a radioactive substance to undergo radioactive decay.

inhibitor substance that reacts competitively with one of the reactants and, as a consequence, decreases the reaction rate.

initial rate of reaction rate of a reaction immediately after the reactants are brought together.

instantaneous rate of reaction exact rate of a reaction at some precise point in the reaction. It is obtained from the slope of a tangent line to a concentration-time graph.

integrated rate equation relates concentration of a reactant (or product) to the elapsed time from the beginning of the reaction. The equation takes different forms depending on the order of the reaction.

intermediate species that appears in the mechanism of the reaction (*i.e.*, the elementary steps), but not in the overall balanced equation. This is because it is produced in one elementary step and then used up in a succeeding step.

kinetic molecular theory (see collision theory)

kinetic energy in chemical terms, this is the energy of motion. Kinetic energy is related to, but not the same as, the speed of the reacting species.

method of initial rates process that establishes the order of a reaction from the initial rates of a reaction by varying the initial concentration of one reactant while initial concentrations of all other reactants are held constant.

order of reaction relates to the exponents of the concentration terms in the rate law for a chemical reaction.

potential energy energy of position, or stored energy.

potential energy diagram graphical device to show the energy relationships in a chemical reaction as the reaction progresses; reaction profile.

rate constant proportionality constant that permits the rate of a reaction to be related mathematically to concentrations of the reacting species.

rate law expression relating the rate of reaction to the rate constant and to concentrations of the reacting species.

rate-determining step in a reaction mechanism, the elementary process that is instrumental in establishing the rate of the overall reaction, usually because it is the slowest step in the mechanism.

reaction mechanism pathway by which a chemical reaction occurs. Almost all reaction mechanisms are believed to occur by one or more two-body collisions (elementary steps).

reaction rate change in concentration of a reactant (decreasing) or a product (increasing) with time.

second-order reaction reaction whose rate depends on reactant concentration raised to the second power, or on the concentrations of two different reactants, each raised to the first power.

side reaction (concurrent reaction) reaction that occurs at the same time as the main reaction in a chemical process. Usually the existence of a side reaction decreases the yield of the desired product(s).

termolecular reaction elementary reaction that involves three molecules simultaneously colliding.

transition state theory describes a chemical reaction through a hypothetical intermediate species called an activated complex. The activated complex dissociates either back into the original reactant molecules or into product molecules.

unimolecular reaction elementary reaction involving only one reactant molecule.

zero-order reaction reaction that proceeds at a rate that is independent of reactant concentration. The concentration-term exponent(s) in the rate law is (are) equal to zero.

Common Student Misconceptions

1. **"The mechanism of a reaction can be deduced from the overall balanced chemical equation."**

   Actually, the overall equation rarely relates to the mechanism, which usually is a series of unimolecular and/or bimolecular steps. If the reaction occurs in one step, then the mechanism can be deduced from the balanced chemical equation.

2. **"A catalyst produces higher a yield of product."**

   A catalyst may produce product faster, but the theoretical amount will be the same.

3. **"A catalyst increases the rate of the forward reaction by lowering the activation energy."**

   The catalyst provides a new path with an accompanying lower energy of activation, but the *old* path with its higher activation energy also still remains.

4. **"One can increase the rate of reaction by vigorous stirring, which results in more energetic collisions."**

   Stirring or shaking is needed for homogeneous distribution of reactant species, not to significantly increase the average energy of the molecules.

5. **"If experimental data are fully consistent with a proposed mechanism designed to explain what occurs at the molecular level, that mechanism must be the correct one."**

   More than one mechanism may be consistent with the same set of experimental data.

6. **"A catalyst is a species that increases the rate of reaction but it does not change or participate in the reaction."**

   Why would a catalyst need to be present if it didn't do anything? The catalyst is directly involved in the reaction, often being used up and then regenerated at the end of the reaction. A catalyst appears in a reaction mechanism first in one step as a reactant and, in a later step, as a product.

7. **"A negative catalyst is a species that slows the rate of a reaction by raising the activation energy."**

   First, there is no such thing as a "negative" catalyst. Even if a species were introduced that provided a path with higher activation energy, the old path with its (relatively) lower activation energy would still be available. Actually the term "inhibitor" is more appropriate. An inhibitor is a species that slows down the rate by reacting with a reactant to form new products. This competition lowers reactant concentration, which, in turn, slows the reaction rate.

8. **"Exothermic reactions are always spontaneous, while endothermic reactions are very difficult to initiate."**

   The difficulty/ease of initiating a reaction depends on the magnitude of the activation energy. Equilibrium concentrations of products are determined by ΔG of the reaction. Products are increasingly favored as ΔG becomes more negative. Since $\Delta G = \Delta H - T\Delta S$, a negative value of ΔH and/or a positive value of ΔS drives the reaction forward. Kinetics and equilibrium thermodynamics are two independent factors affecting chemical change. Some endothermic ($\Delta H > 0$) changes are spontaneous, *e.g.*, dissolving of sodium acetate or ammonium chloride in water, while some exothermic reactions are nonspontaneous. For example, the condensation of steam to liquid water—an exothermic reaction—is nonspontaneous above 100 °C at 1 atm pressure. Also, stretching a rubber band (obviously not spontaneous) is an exothermic reaction, while releasing a stretched rubber band (obviously spontaneous) is endothermic.

Pictures in the Mind

1. Molecules must (a) collide with (b) sufficient energies in order to react and must collide with an appropriate (c) collision geometry. (a) When two slow-moving molecules collide, their electron clouds cannot interpenetrate much—they just bounce off each other, chemically unchanged. (b) When fast-moving molecules collide, atoms approach each other much more closely as their clouds interpenetrate. This can lead to bond-making and bond-breaking. The net change is $AB + C \rightarrow A + BC$.

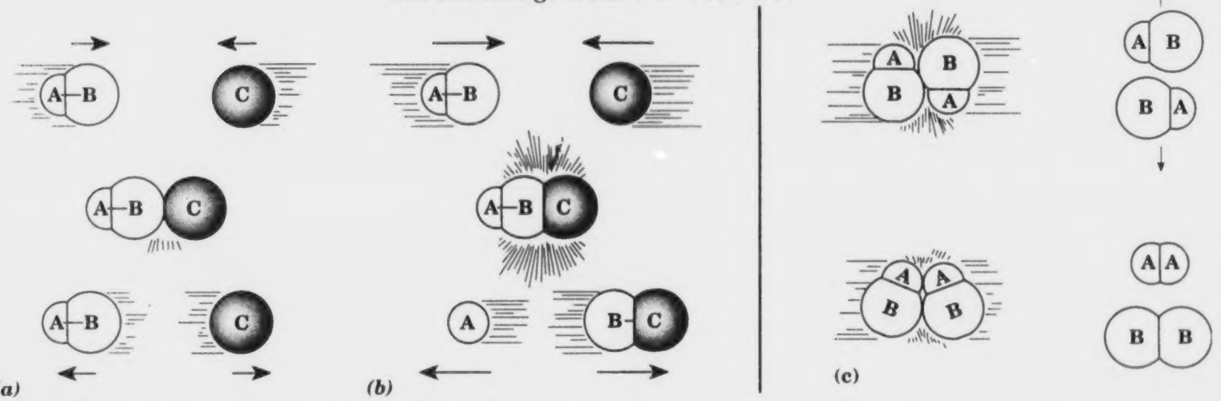


Figure 15. Effects of molecular speed on reaction.

2. How concentration and reaction rates are related. The number of collisions per second depends upon the concentration of each reactant.

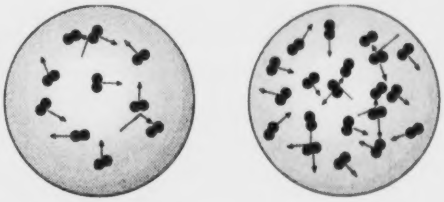
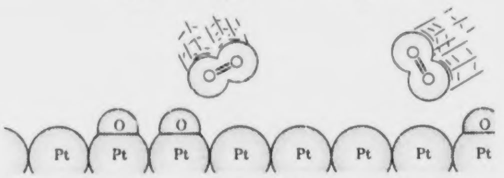


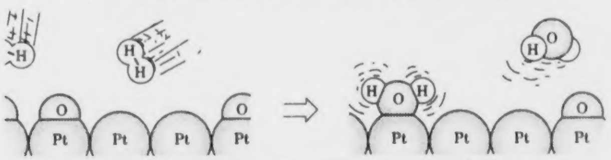
Figure 16. Effect of concentration on reaction rates.

3. A possible mechanism for the platinum catalyzed reaction:

$$2H_2(g) + O_2(g) \overset{Pt}{\rightarrow} 2H_2O(l).$$



(a) Oxygen molecules react with the platinum surface forming atoms.



(b) Hydrogen molecules react with oxygen atoms forming water.

Figure 17. Reaction catalysis.

4. Simple reaction mechanisms for exothermic bimolecular reactions.

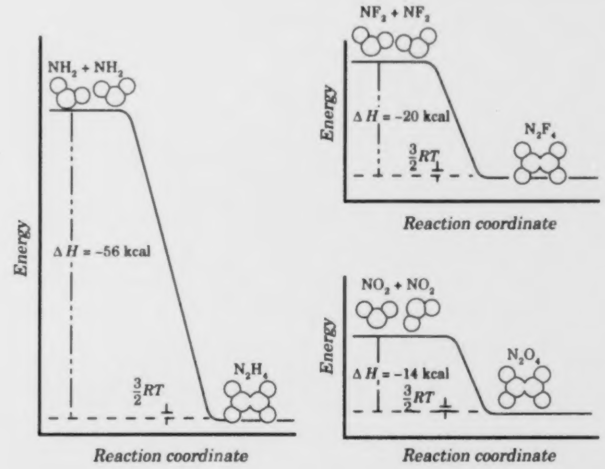

NH₂ + NH₂
Energy
ΔH = −56 kcal
3/2 RT
N₂H₄
Reaction coordinate
NF₂ + NF₂
Energy
ΔH = −20 kcal
3/2 RT
N₂F₄
Reaction coordinate
NO₂ + NO₂
Energy
ΔH = −14 kcal
3/2 RT
N₂O₄
Reaction coordinate

Figure 18. Three similar exothermic reactions.

5. Unimolecular reaction mechanism.

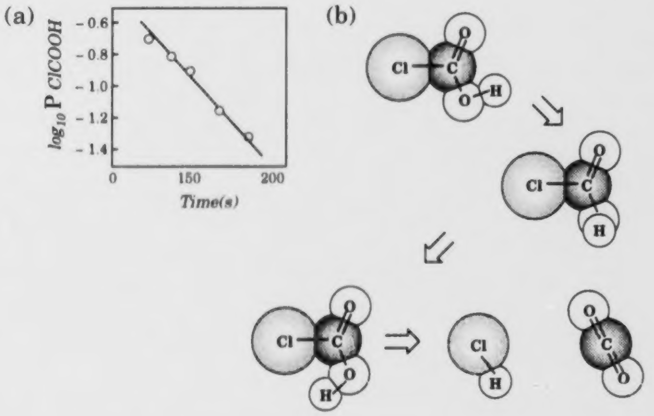

(a)
log₁₀ P ClCOOH
−0.6
−0.8
−1.0
−1.2
−1.4
0
150
200
Time(s)
(b)
Cl
C
O
O
H
Cl
H

Figure 19. Decomposition of chloroformic acid (ClCOOH): (a) graphical and (b) mechanism.

6. Catalyzed and uncatalyzed mechanism of formic acid decomposition.

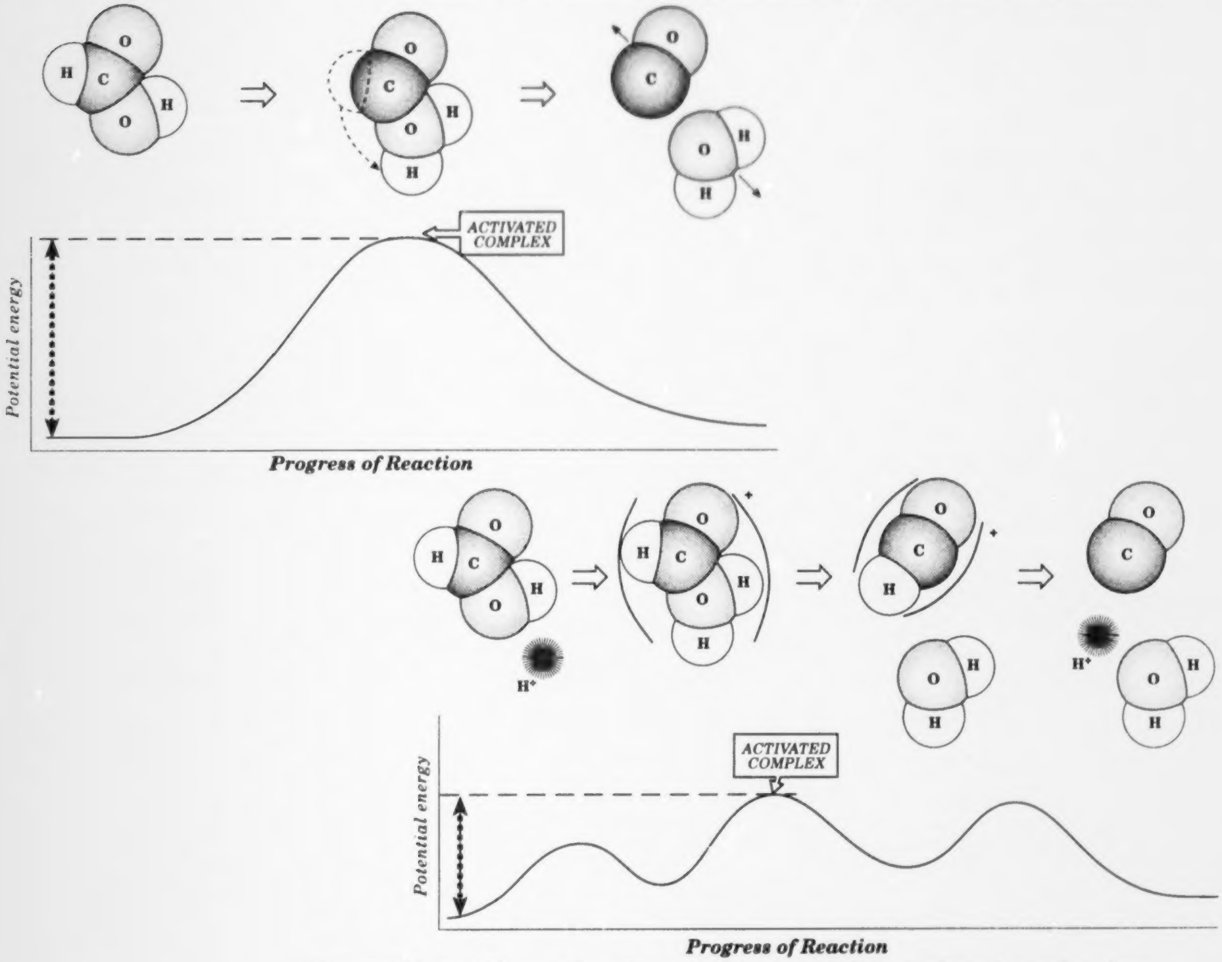

H
C
O
O
H
ACTIVATED COMPLEX
Potential energy
Progress of Reaction
H⁺
+
ACTIVATED COMPLEX
Potential energy
Progress of Reaction

Figure 20. (a) Uncatalyzed mechanism and (b) catalyzed mechanism.

Problem Solving

1. It normally takes 20 min at 100 °C to insure that an egg is fully hard-boiled (hard cooked). If the reaction involved—the denaturation of egg albumin—doubles in rate for each 10 °C increase in temperature, how long would it take to "hard cook" an egg on a mountain plateau at each of these locations?

Location	Approximate Altitude (km)	Water's B.P. (°C)	Cooking Time (min)
(a) Mt. Montcalm, Spain	3	90	
(b) Mt. McKinley	6	80	
(c) Mt. Everest	9	70	

[(a) 40 min, (b) 80 min, (c) 160 min]

Figure 21. Altitude effects on cooking time.

2. An airline flight of 1200 miles from Philadelphia to Miami takes 2 hours and 10 min. What is the average speed in: (a) miles/h; (b) miles/min; (c) km/h; and (d) cm/s. How would the rate change with: (e) a tail wind; (f) a head wind; (g) a rough rain storm. *[(a) 550, (b) 9.2, (c) 890, (d) 2.5 x 10^4, (e) increase, (f) decrease, (g) decrease]*

3. You leave your home to go to the airport from a distance 100 km away by car. You park your car at the airport and go through the normal stages following check in. Upon landing at your destination, you rent a car and drive to your hotel. List the different steps from your home to your destination and indicate the relative rates. Discuss which step might be the rate-determining step.

   Drive → Park → Check-in → Walk to gate →
   Gate check → Plane ride → Rent car → Drive to hotel

4. For the overall reaction, $NO_2(g) + CO(g) \rightarrow NO(g) + CO_2(g)$, two mechanisms have been proposed.

   I. $NO_2 + NO_2 \rightarrow N_2O_4$ Slow
   $N_2O_4 + CO \rightarrow NO + CO_2 + NO_2$ Fast
   and II. $NO_2 + NO_2 \rightarrow NO_3 + NO$ Slow
   $NO_3 + CO \rightarrow NO_2 + CO_2$ Fast

   Experimental facts include: Rate = $k[NO_2]^2$, and the detection of NO_3. Discuss the two mechanisms in terms of the experimental data. *[On the basis of the slow rate determining step both given mechanisms would agree with the rate expression but only mechanism II would have an NO_3 intermediate.]*

5. The activation energy reaction profile diagram (potential energy diagram) for the three-step mechanism for the overall reaction $H_2(g) + I_2(g) \rightarrow 2\,HI(g)$ is shown in Figure 22.

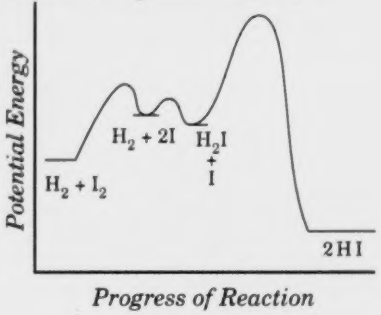


Figure 22. Activation energy for a three-step reaction.

Complete the following table:

	Exothermic or Endothermic	Relative Rate	Equation for Reaction
Step 1			
Step 2			
Step 3			

Answers:

	Exothermic or Endothermic	*Relative Rate*	*Equation for Reaction*
Step 1	*Endothermic*	*Fast*	$I_2(g) \rightarrow 2I(g)$
Step 2	*Exothermic*	*Fastest*	$I(g) + H_2(g) \rightarrow H_2I(g)$
Step 3	*Exothermic*	*Slow*	$H_2I(g) + I(g) \rightarrow 2HI(g)$

HISTORY: ON THE HUMAN SIDE

1. Early developments in chemical reaction rate theory were, and still are, closely tied to catalysis. Jons Jakob Berzelius first named and defined catalysts in 1836, drawing on earlier work of Johann Wolfgang Döbereiner. Döbereiner had discovered in 1823 that hydrogen mixed with air would ignite on contact with a platinum sponge.

   The direct oxidation of sulfur dioxide with air, using a platinum catalyst (forerunner of the contact process), was first noted by Teregrine Phillips in 1831 and restudied by Clemens Winkler in 1875. The first successful commercial use of this contact process did not occur until 1901.

   The goal of many chemists around 1900 was the synthesis of ammonia from hydrogen and nitrogen. Fritz Haber, using iron as a catalyst, saw industry first use his process in 1913. Since conversion of ammonia to nitric acid was already commercial at that time, Haber's process meant that Germany was no longer dependent on Chilean saltpeter for nitrate-based explosives—a great advantage to the German government during World War I.

   The 1930s saw a spurt in technological applications of catalysis, mainly in the field of petroleum conversion to high-octane gasoline. Petroleum products were later expanded to include plastics, synthetic lubricating oils, and plasticizers.

   [The preceding was condensed from the following: Haensel and Burwell. (1971, December). Catalysis. *Scientific American, 255*(6).]

2. Chemical reaction rates were traditionally studied by mixing several solutions and observing resultant changes. Since such mixing takes several seconds, study was limited to reactions with rates no faster than this.

   In 1923, Hartridge and Roughton devised a way to study much faster reactions by mechanical mixing of reactants and observing resulting color changes in a long tube. Their method allowed study within a few milliseconds of a reaction's starting time. Because many chemical reactions occur rapidly, development of such methods was essential to further study.

   Flash photolysis, in which a brief, strong flash of light is used to start a reaction (replacing the use of a steady light source for this purpose), was devised by George Porter and R. G. W. Norrish in 1949. It was subsequently used by Gerhard Herzberg and D. A. Ramsay, and allowed study of reactions whose intermediate products formed a few milliseconds after their start.

   In 1954 Manfred Eigen devised a method to permit measurement of nanosecond (billionths of a second) intervals. He studied the sudden upsetting of a system at equilibrium by pressure or temperature changes. This method, known as "relaxation," won him the Nobel prize in 1967.

   [The preceding was condensed from the following: Grossweiner, L. (1960, May). Flash photolysis. *Scientific American, 202*(5), 135; and Faller, L. (1969, May). Relaxation methods in chemistry. *Scientific American, 220*(5), 30.]

3. With development of laser technology, chemists could measure effects within short time intervals. Nesbit and Leone, using microsecond laser pulses, were able to measure photodissociation times in the microsecond range for $S_2Cl_2 \rightarrow S_2Cl + Cl$ in the early 1980s.

   Schere and colleagues in 1985 used even shorter laser pulses, 400 femtoseconds long, to measure the lifetime of excited ICN molecules. These dissociate according to the equation $ICN \rightarrow CN + I$. The time required for this dissociation was found to be about 600 femtoseconds. One femtosecond is one quadrillionth of a second, 0.000 000 000 000 001 s, or 1×10^{-15} s.

   Similar studies, using 400 femtosecond laser pulses, were undertaken by Knee and co-workers in the mid-1980s. They measured the time required to break $I\text{–}CH_3$ bonds to within 400 femtoseconds after excitation.

   (The preceding was condensed from the following: Kovalenko and Leone. (1988, August). Innovative laser techniques in chemical kinetics. *Journal of Chemical Education, 65*(8), 681.)

HUMOR: ON THE FUN SIDE

1. Sign on a bumper sticker: Warning, I have a very low activation energy!
2. Messages on T-shirts:
   a. I'm a Catalyst, Let's Go!
   b. Chemists are quick to react.
   c. Old chemists never die: They just fail to react.
3. **An Ode to Kinetics**

In this lab we learned a lot
  Of rates of a reaction:
The effects thereon of concentration
  In speeding or slowing the action.

All of the thiosulfate consumed
  Before we stopped the clock;
It used the iodine produced.
  I spilled some on my smock.

The peroxodisulfate reaction was
  The slower of the pair,
And thus the rate-determiner.
  God thought it only fair.

And thus by finding out the rate
  Which consumed S_2O_3;
We discover the rate of S_2O_8,
  And write it log base e.

By varying the concentrations
  Of our renowned reagents,
We vary the rate; but also by
  Our catalyzing agent.

We love you H_2SO_4!
  You lower energy
Of activation for our reaction,
  So it speeds up. Oh Gee!

Through careful calculations we
  Are able to obtain
The orders of reactants and
  From that, the constant k.

Our errors come from measurements
  And ill-begotten data;
From numbers gathered from classmates
  Of varied IQ strata.

So let it be heard far and wide,
  In every tavern, chamber, and hall;
'Tis better to have mixed and spilt,
  Than never to have mixed at all.

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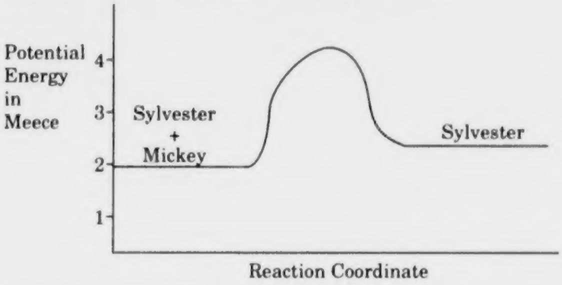
4. **Sylvester the Cat and Mickey Mouse**



Consider this reaction:

Sylvester the Cat + Mickey Mouse ⟶ Sylvester the Cat

a. Is this an endothermic or exothermic process?

b. What is the ΔH value for this system?

c. Where would the activated complex be located on this graph?

d. Why would concentration changes have no effect on the rate of this reaction?

e. What would the activated complex look like?

f. From a practical viewpoint is this reaction reversible? Explain.

g. From a practical viewpoint, what is the most favored equilibrium state? Explain.

CHEM 13 NEWS, September 1979, p. 7

5. Word Search (see *Appendix* for master copy)

H F F V W E ~~P O T E N T I A L~~ M X
B O T P S Y F T F G N W N Z G D K
C ~~M E C H A N I S M~~ U Z D F F X W
L P T G F X V I L L O F Y T Z E B
S S H V F Y X K O F R C F M X W X
C C E Q Z K L D G S L P G D E O G
I A Y ~~C O N C E N T R A T I O N~~ A
T T J M M R M B U F G M N X R P M
E A F ~~P R O D U C T~~ N T J B U S P
N L Y C U O R E K C J X Q Z N F J
I Y N G C Y B T R B F P W B Z A W
K S N U F A U F ~~T N A T C A E R~~ W
N T P F P M O Z L X L G V E U X I

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. mol/L or weight %, for example.
2. Time required for one-half of a given amount of reactant to be consumed in a chemical reaction (2 words).
3. Species used up in a chemical reaction.
4. Number, usually integral, in the rate expression for a chemical reaction.

   5. Chemical reaction pathway.
   6. Adjective for stored energy.
   7. Study of speeds or rates at which chemical reactions occur.
   8. Biological catalyst.
   9. Species formed in a chemical reaction.
   10. Substance that provides a different pathway of lower energy for a chemical reaction.

   Answers: 1. CONCENTRATION 2. HALF LIFE 3. REACTANT 4. ORDER 5. MECHANISM 6. POTENTIAL 7. KINETICS 8. ENZYME 9. PRODUCT 10. CATALYST

6. See cartoons at end of module.

MEDIA

1. *TheWorld of Chemistry* videotape "Number 13: The Driving Forces," explains why reactions occur and discusses spontaneity and nonspontaneity. World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.
2. The World of Chemistry (High School Version), available from WINGS for Learning/SUNBURST, 101 Castleton Street, Pleasantville, NY 10570; (914) 747-3310; (800) 321-7511; (914) 747-4109 (FAX).
   a. *Chemical Kinetics*. This 15 min. video discusses why some reactions take place rapidly while others are slow and how reaction rates can be altered. A good combination of theory, demonstrations and industry/society applications.
   b. *Catalysts*. This 15 min. video shows how catalysts function on a molecular level.
3. *Reaction Rates and Equilibrium*. Available from Coronet/MTI Film and Video, 108 Wilmot Road, Deerfield, IL 60015; (800) 621-2131; (708) 940-3640 (FAX). This 20 min. video explores fundamental factors that influence the rates at which chemical reactions take place. It also explores equilibrium, how it can be upset, and its role in the chemical industry.
4. CHEM Study films and videos available from Ward's Natural Science Establishment, Inc., P.O. Box 92912, Rochester, NY 14692-9012; (800) 962-2660.
   a. *Catalysis*
   b. *An Introduction to Chemical Kinetics*
   c. *Molecular Motion*
   d. *Reaction Rates and Equilibrium*
   e. *Mechanism of an Organic Reaction*
5. Software published by *JCE*: Software, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
   a. *Reaction Dynamics*, by Daniel Lacks. Vol. IV C, No. 2, for the Apple Macintosh.
   b. *Animation of a Atom-Molecule Chemical Reaction*, by Douglas Kutz. Vol. III C, No. 2, for the Apple Macintosh.
   c. *The Computer-Based Laboratory*, by Daniel Krause. Vol. I A, No. 2, for the Apple II computer.

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6. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).
    - a. For the Apple II computer running on ProDOS: AR 1201
    - b. For the Apple II computer: AP 603 (includes animated program showing the chain reaction involving chlorination of methane; useful classroom tool to show a reaction mechanism, complete with initiation, propagation and termination steps), AP 607 (reaction rates laboratory simulation), AP 702, AP 1001 (Decay: Drill program on radioactive decay as a first-order process), AP 1201, as well as Laboratory Modules LM 002, LM 003, LM 004, LM 005, LM 006, LM 012, LM 015, LM 016, LM 021 for the Project SERAPHIM Blocktronic colorimeter.
    - c. For IBM PCs and PC-compatibles: PC 4201, as well as Laboratory Modules LM 002, LM 003, LM 004, LM 005, LM 006, LM 012, LM 015, LM 016, LM 022 for the Project SERAPHIM Blocktronic colorimeter.
7. *Chemical Principles for the Introductory Laboratory*. Available from Conduit, University of Iowa, Oakdale Campus, Iowa City, IA 52242; (800) 365-9774. "First Order Rate of Reaction" is one of five programs on this disk. The program asks students to select a temperature and starting concentration for one of three limiting reagents. It then displays a table of concentration *vs*. time data, plots values from the table and helps students see the effect of changes in temperature and concentration on reaction rates. Accurate record-keeping is essential, since plots cannot be saved.
8. Videodiscs published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
    - a. "A Catalyst on the Surface" and "The Boat and Dock Model," two chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc II (double sided, 60 min.), Special Issue 4.
    - b. "Activation Energy," a chapter on *The World of Chemistry: Selected Demonstrations and Animations*: Disc I (double sided, 60 min.), Special Issue 3.
    - c. "Nucleophilic Substitution Reactions," a chapter on the videodisc *Demonstrations in Organic Chemistry* (double sided, 60 min.), Special Issue 6.
    - d. *The Periodic Table Videodisc* (single side, 30 min.). Special Issue 1. The section on alkali metal reactivity can be used to show how the nature of the reactant can affect reaction rate. Sequences showing alkali metals reacting with water illustrate this concept very nicely.

EQUIPMENT AND INSTRUMENTATION

1. *Laboratory Activity 1* involving the iodine-clock reaction may lend itself to using the *Spectronic-20* or the *ChemAnal* system to determine the formation of the blue color, although the color is so obvious that this addition to the procedure may seem superfluous.
2. The Blocktronic colorimeter that attaches to the Apple computer may also be used in the iodine-clock reaction. This could be programmed to time the reaction automatically by initializing the timer when the two solutions first come in contact and then stopping the timer with the first appearance of blue color (represented by a sudden decrease in transmittance of light through the solution in the sample tube). The computer could then print out the reaction time. Directions for making the Blocktronic (LM 002) are available from Project SERAPHIM (address given in *Media* section). Universal adapter box for the Apple (LM 021) or for the IBM PC (LM 022) are essential for connecting the Blocktronic colorimeter to the computer of choice.

Links/Connections

WITHIN CHEMISTRY

1. The kinetic theory explains gas behavior in terms of a model involving rapidly moving molecules colliding with container walls and with each other. Mixtures of two or more gases, each exerting a partial pressure, but not reacting with each other, exemplify ineffective collisions. In both cases, the molecules rebound completely unchanged. A common example is the absence of chemical reaction, at normal temperatures, between gaseous oxygen and nitrogen molecules colliding constantly in ordinary air.
2. The reactivity of alkali metals with water for example, in contrast with less reactive metals such as iron and copper, offer striking examples of how the nature of reactants can affect reaction rate.
3. Radioactive decay is an excellent example of a first order reaction. The shorter the half-life of an isotope (nuclide) the greater the rate at which the atom decays. Examples,

$^{226}Rb \rightarrow {}^{222}Rb + {}^{4}He$	$t_{1/2} = 1600$ y	Slow
$^{238}U \rightarrow {}^{234}Th + {}^{4}He$	$t_{1/2} = 4.5 \times 10^9$ y	Very slow
$^{211}At \rightarrow {}^{207}Bi + {}^{4}He$	$t_{1/2} = 7.5$ h	Fast

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

1. **Physics**. Momentum and conservation of linear momentum involve particles colliding both elastically and inelastically. Elastic collisions are examples of ineffective collisions, resulting in unchanged particles rebounding. Inelastic collisions, such as those involved in auto accidents, exemplify effective collisions, since they produce different products.
2. **Earth Science**. Erosion involves slow reactions, while volcanic eruptions are extremely rapid.
3. **Biology**
    a. The rate of photosynthesis in plants depends, in part, on temperature.
    b. The rate of production of carbon dioxide by yeast depends both on temperature and yeast particle size. Powdered yeast produces CO_2 faster than does lump yeast.
    c. Some warm-blooded animals' metabolic reactions slow during winter hibernation, and speed up during warmer spring and summer months.
    d. Cold-blooded animals become dormant in cold weather.
    e. Application of cold temperatures to induce hypothermia has been used in certain medical procedures, as well as in science fiction accounts of astronauts put to sleep during long space journeys, so that low temperatures slow their aging processes.

TO THE CONTEMPORARY WORLD

Personal

1. Chemical reactions in food can be slowed to prevent spoilage by storing foods at lower temperatures in refrigerators or freezers, or by using chemical inhibitors, *e.g.*, calcium propionate in bread. Reactions can be speeded up by raising temperature, *e.g.*, by cooking.

2. Campfires or home fireplace fires are started with small kindling rather than large logs to increase the surface area of wood in contact with air. For example, compare the total surface area of a 1-L cube of wood, 10 cm on an edge, with the same 1-L cube cut into 1000 1-cm^3 cubes.)

Societal

1. Stratospheric ozone depletion by chlorofluorocarbons (CFCs) involves catalytic cycles in which chlorine atoms from CFCs are first consumed, and later regenerated. Thus the catalyst, chlorine, is not consumed. Several key reaction rates had to be measured before predictions and/or explanations of ozone loss could be made.
2. Explosions of fine dust particles in confined places, as in grain elevators or saw mills, are due to large surface areas of combustible material being exposed to oxygen molecules in air, allowing for rapid combustion if ignited. Similarly, explosions of gases such as methane in coal mines are due to oxygen molecules in the air intermingling with the combustible gas molecules.
3. Proposed nuclear waste storage sites often experience public opposition due to the slow decay of certain fission fragments with very long half-lives. Such fragments will remain potentially hazardous for hundreds of years or more.

Extensions

1. Students can explore one of the two mechanical analogies provided in this module under *Metaphors and Analogies*.
2. Student reports:
    a. Review of modern kinetics—a *Scientific American* literature search.
    b. Rates of reaction in everyday life.
    c. Kinetics concepts — a review of the *Journal of Chemical Education*, 1970-present.
3. Discuss the use of radioactive isotopes as tracers. In the context of half-lives, discuss health radiation dangers associated with applications for these first-order reactions.
4. Explore the mechanism involved in destruction of the ozone layer from chlorofluorocarbons, supersonic transports, and other sources in terms of chain-reaction mechanisms. Suggested reading for research: Brune, Anderson *et al.* (1991, May 31). The potential for ozone depletion in the arctic polar stratosphere. *Science, 252*(5010), 1260-1266.
5. Complete a library investigation of the mechanism of hydrogenation of vegetable oils, such as safflower, corn, canola, cottonseed, palm and coconut oils, using finely divided nickel and high temperatures to form fats. Discuss nutritional aspects of saturated, mono- and polyunsaturated fats.
6. The mystery flask experiment (see *Appendix*) is ideal for individual or small-group class laboratory projects.

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Module developed by Herbert Bassow, William Bleam, and Arthur C. Breyer, the Pennsylvania team.

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Hammes, G. G. (1978). *Principles of chemical kinetics.* New York, NY: Academic Press.

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Demonstrations on kinetics are on pp. 69-83.

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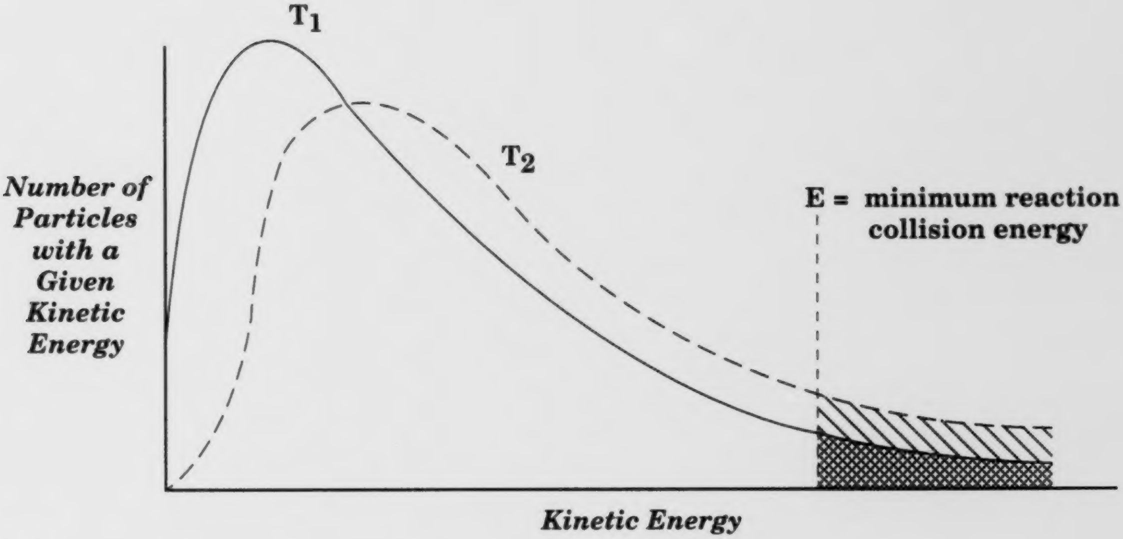
Appendix

- **Transparency Masters**
  1. Effect of Temperature on Kinetic Energy Distribution
  2. The Mechanism and Potential Energy Diagrams for the Reaction
  3. Rate-Determining Steps in Three Reactions
  4. Set-Up for Hydrogen Peroxide Decomposition
  5. Word Search
- **Activity 3: The Mystery Flask—An Introduction to Reaction Rates**
- **Humor**

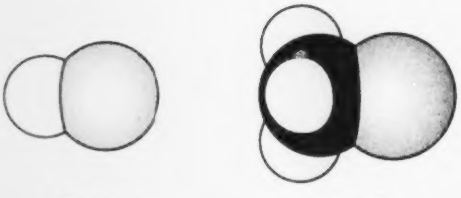
Effect of Temperature on Kinetic Energy Distribution

T2 Greater Than T1



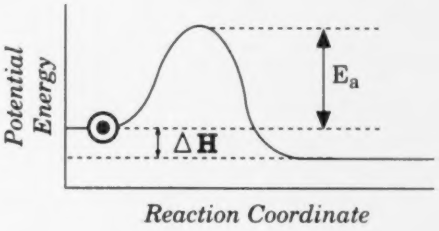
The Mechanism and Potential Energy Diagrams for the Reaction

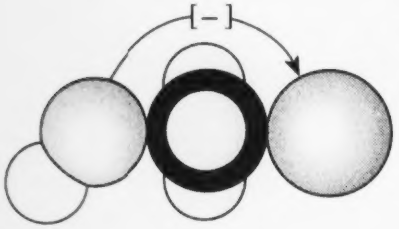
$CH_3Br + OH^-(aq) \longrightarrow CH_3OH + Br^-(aq)$



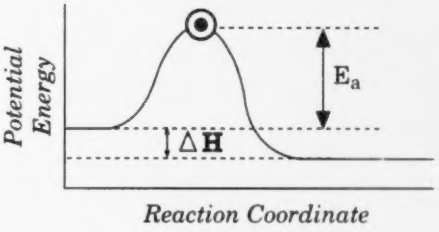
$OH^-(aq)$ + CH_3Br

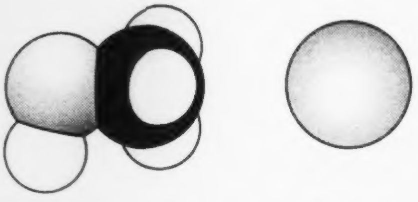
Reactants Approaching


Potential Energy
E_a
ΔH
Reaction Coordinate



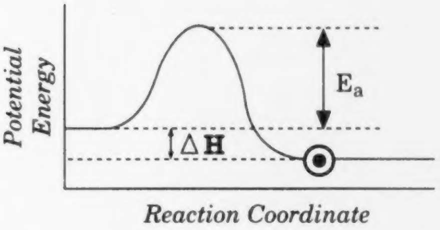
Activated Complex


Potential Energy
E_a
ΔH
Reaction Coordinate



$HOCH_3$ + $Br^-(aq)$

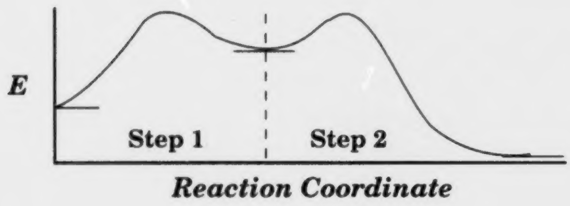
Products Separating


Potential Energy
E_a
ΔH
Reaction Coordinate

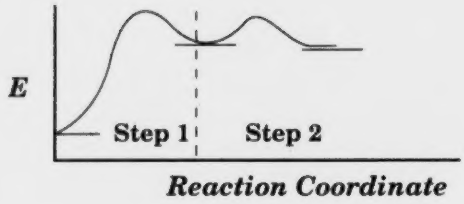
E_a = Activation Energy; ΔH = Enthalpy Change

Rate-Determining Steps in Three Reactions

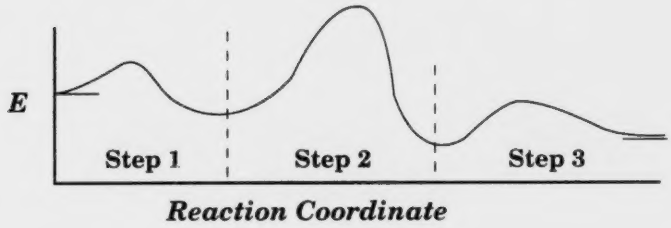
Ⓐ


E
Step 1
Step 2
Reaction Coordinate

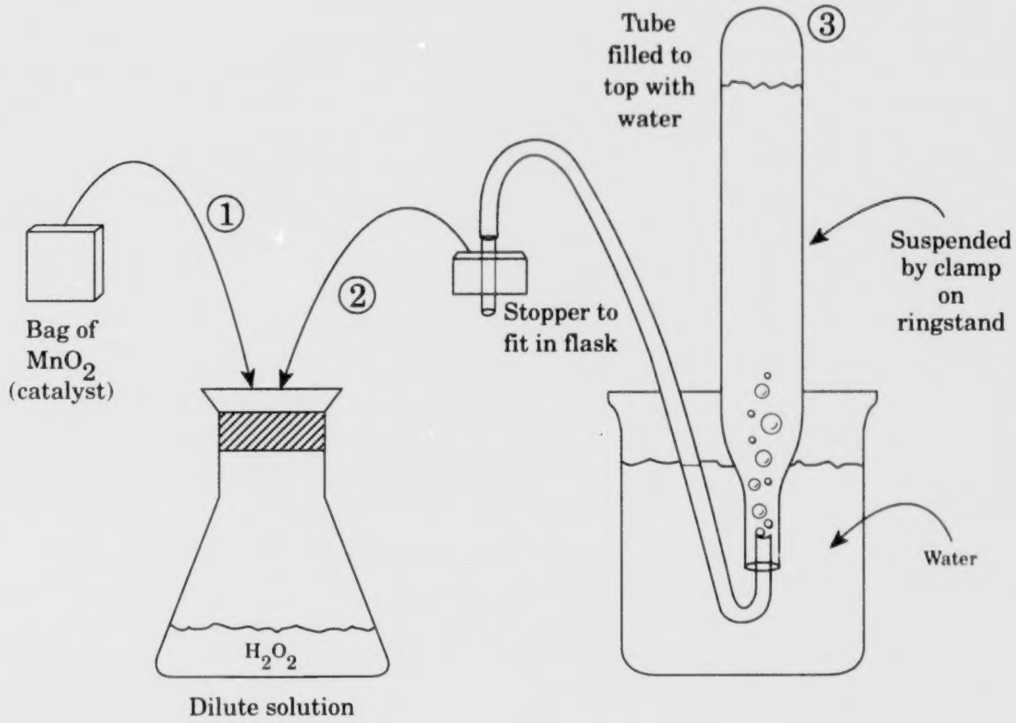
Ⓑ


E
Step 1
Step 2
Reaction Coordinate

Ⓒ


E
Step 1
Step 2
Step 3
Reaction Coordinate

Set-Up for Hydrogen Peroxide Decomposition


Tube filled to top with water
③
①
②
Bag of MnO$_2$ (catalyst)
Stopper to fit in flask
Suspended by clamp on ringstand
Water
H$_2$O$_2$
Dilute solution

Word Search

```
H F F V W E P O T E N T I A L M X
B O T P S Y F T F G N W N Z G D K
C M E C H A N I S M U Z D F F X W
L P T G F X V I L L O F Y T Z E B
S S H V F Y X K O F R C F M X W X
C C E Q Z K L D G S L P G D E O G
I A Y C O N C E N T R A T I O N A
T T J M M R M B U F G M H X R P M
E A F P R O D U C T N T J B U S P
N L Y C U O R E K C J X Q Z N F J
I Y N G C Y B T R B F P W B Z A W
K S N U F A U F T N A T C A E R W
N T P F P M O Z L X L G V E U X I
```

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. mol/L or weight %, for example.
2. Time required for one-half of a given amount of reactant to be consumed in a chemical reaction (2 words).
3. Species used up in a chemical reaction.
4. Number, usually integral, in the rate expression for a chemical reaction.
5. Chemical reaction pathway.
6. Adjective for stored energy.
7. Study of speeds or rates at which chemical reactions occur.
8. Biological catalyst.
9. Species formed in a chemical reaction.
10. Substance that provides a different pathway of lower energy for a chemical reaction.

LABORATORY ACTIVITY: STUDENT VERSION

Activity 3: The Mystery Flask—An Introduction to Reaction Rates

Introduction

The production of important industrial chemicals, such as ammonia, $NH_3(g)$, requires a set of conditions that give the highest possible yield, while, at the same time, maintaining a rapid controllable rate of production. The rate, or speed, of a chemical reaction is measured by the decrease of the concentration of any reactant, or the increase in the concentration of any product, in a given interval of time. The study of the rates of chemical reactions is termed *chemical kinetics*, or *reaction kinetics*. On a microscopic level, for two molecules to react they must (a) collide, (b) collide with sufficient energy, and (c) collide with a suitable spatial orientation. As a consequence, the rate of a reaction can be changed by any factors that alter the number of collisions, the average collisional energy, or the collisional geometry. In this activity, you will investigate a number of experimental variables to determine how they affect the rate of a reaction.

Purpose

To explore the factors that affect the rates of chemical reactions through the scientific investigation of a fascinating chemical change and to begin to gain an understanding of how and why chemical substances change on a molecular level (mechanism of a reaction).

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Sodium hydroxide, NaOH, is a very corrosive reagent. If any solution is spilled on your desk or hands, rinse thoroughly with water followed by a rinse with dilute acetic acid (vinegar).
3. Dispose of substances as your teacher directs.
4. Thoroughly wash your hands before leaving the laboratory.

Procedure

Part I of this laboratory activity will be done with groups of two to four students. This will be followed by a class discussion, during which you will be shown the results of two other experiments. Your teacher will then assign groups of two to three students to study ***one or more of the experimental variables in* Parts II *to* VII**. Make careful observations and try visualizing what might be occurring on a molecular level. *NOTE: Your teacher may wish to carry out this laboratory activity as a small group science project or as a series of student and/or teacher-led demonstrations correlated with extensive discussion. In the past, some teachers have used the initial experiment simply as an experiment in which students collect their individual observations,* i.e., *an experiment on observation.*

Part I. Background Experience

1. Your teacher will provide you with a stoppered flask containing a solution marked X. Observe carefully, and describe the initial appearance of the solution.
2. Hold the stopper and flask tightly and shake vigorously five times. Carefully observe and record your observations.

3. Repeat the shaking a few times, varying the number of shakes. Where does the color persist longest in the solution?
4. The Solution X contains NaOH(aq), dextrose (glucose), and a few drops of methylene blue dye. Two other solutions are available in the laboratory. Y, a system that has stood overnight, and Z, a system without the methylene blue, which also has stood overnight. Describe the behavior of Y and Z, and compare with X. Can you arrive at any conclusions?
5. You will now study the experimental variables. Your teacher will assign you one or more of the following studies to carry out. Eventually, the class data will be collected and each student given a copy. A stopwatch or clock is required for these experiments.

Part II. Variation of the Number of Shakes

1. Take a clean 250-mL flask and stopper. Measure out 50.0 mL 1.5 M NaOH(aq) and 50.0 mL 0.133 M aqueous glucose (dextrose) and add them to the flask. Add 3 drops of 0.1% methylene blue dye solution, record the temperature, and tightly stopper the flask.
2. To study the effect of shaking, you must develop your own "standard" shake. A sharp vertical shake will probably work best. Whatever shake you use, apply it consistently. Make sure that the rubber stopper is firmly in place and hold it tightly while shaking. Shake once and record with your timer the time it takes for the color to disappear from the bulk of the solution. (Disregard any tiny specks of blue near the top on the walls of the flask.)
3. Repeat the timing for each of the number of shakes indicated on the record sheet. Remove the rubber stopper before each experimental run and then replace it immediately prior to the next run. Data columns are provided for additional class data.
4. Plot the number of shakes (vertical y-axis) *versus* the time for the color to disappear (horizontal x-axis). Draw the best smooth curve through the experimental points. What conclusions can you draw?

Part III. Variation of the Concentration of Dextrose

To study the effect of changing any one variable on the rate of a reaction, all the other variables should be kept constant:

Number of shakes	3 or 5 (record choice)
Drops of methylene blue	5 or 7 (record choice)
Temperature	Record
Volume of NaOH(aq)	50.0 mL

1. Vary the concentration of dextrose as indicated on the record sheet. Using your standard shake, perform sufficient experiments to obtain an evenly spaced range of time intervals for the color to disappear.
2. Plot the volume of dextrose (in mL) *vs.* the time in seconds, and draw the best smooth line through the points. What conclusions can you draw?

Part IV. Variation of the Concentration of NaOH

Volume of dextrose	50.0 mL
Number of shakes	3 or 5 (record choice)
Drops of methylene blue	5 or 7 (record choice)
Temperature	Record

1. The concentration of NaOH(aq) will be varied as indicated on the record sheet.
2. Perform enough trials to obtain an evenly spaced range of time intervals for the color to disappear.
3. Plot the volume (in mL) of NaOH(aq) *vs.* the time in seconds; draw the best smooth line through the points. What conclusions can you draw?

Part V. Variation of the Concentration of Methylene Blue

The following variables are kept constant:

Volume of NaOH(aq)	50.0 mL
Volume of dextrose	50.0 mL
Number of shakes	3 or 5 (record choice)
Temperature	Record

1. The concentration (number of drops) of methylene blue will be varied as indicated on the record sheet.
2. Record the time for each trial and plot the number of drops of methylene blue *vs* the time in seconds. What conclusions can you draw?

Part VI. Variation of the Temperature

The following variables are kept constant:

Volume of NaOH(aq)	50.0 mL
Volume of dextrose	50.0 mL
Number of shakes	5 to 10 (record choice)
Drops of methylene blue	5 or 7 (record choice)

1. Add 50.0 mL NaOH(aq) and 50.0 mL aqueous dextrose solution to each of three 250-mL flasks.
2. Place one in an ice bath, one in a water bath at about 45 °C, and allow the third to equilibrate at room temperature. Stopper each tightly.
3. When the solution in the ice bath reaches a temperature of about 10 °C, measure the temperature, and add the selected number of drops of methylene blue, stopper tightly, and shake for the selected number of times.
4. For each run, the temperature should be kept constant by carrying out the reaction quickly, and replacing the flask in the appropriate water bath. Record the time required for the blue color to disappear.
5. Measure the temperature of the solution in the flask kept in the 45 °C water bath, and repeat the laboratory activity.
6. Then repeat the activity for the flask at room temperature, recording the temperature just prior to starting.
7. If time permits, repeat the procedure at several other temperatures.
8. Plot the time *vs.* the temperature, and draw the best smooth curve through the points. What conclusions can you draw?

Part VII. Introduction of Selected Gases into the System

The following variables are kept constant:

Volume of NaOH(aq)	50.0 mL
Volume of dextrose	50.0 mL
Number of shakes	2
Drops of methylene blue	5 or 7 (record choice)
Temperature	Room temperature (record)

1. Add 50.0 mL NaOH(aq) and 50.0 mL aqueous dextrose solution to each of four flasks, and insert stoppers.
2. Bubble $O_2(g)$ slowly into Flask 1 until solution is saturated. Stopper the flask.
3. Repeat the procedure with three other gases: compressed air, natural gas (in the fume hood), $N_2(g)$, $CO_2(g)$, or He(g), as directed by your teacher.
4. Pull the stoppers out just enough to quickly add the required number of drops of methylene blue dye solution to each, and quickly replace the stoppers.
5. Give each solution two shakes and record the time for the blue color to disappear. Note that it is important not to let air enter the solutions after they are saturated with a particular gas. Why? What conclusions can you draw from these experiments?
6. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. Review the data for experiments your group carried out to determine how the variable(s) you studied affected the rate of reaction. If possible, summarize your data and graph, if done, on a transparency sheet. What do you think is the molecular explanation of your results?
2. Your teacher will convene a class discussion section. Each group will present its data and its interpretation to the rest of the class. Your teacher will compile the data on a master list and run off copies for each member of the class. Typical questions for discussion include:

   How was the rate of the reaction affected by:

   a. increasing the concentration of dextrose solution?
   b. increasing the concentration of NaOH(aq)?
   c. increasing the concentration of methylene blue?
   d. increasing the temperature?
   e. increasing the number of shakes?
   f. saturating the solution with selected gases?
3. What does the cyclical appearance and disappearance of the blue color suggest about the role of the methylene blue in the reaction?
4. On the basis of *Parts I* and *II*, why is the number of shakes an important variable? Write down the simple relationship that describes what occurs during the shaking process.
5. The initial color of a freshly prepared solution before the addition of the methylene blue is colorless. After standing overnight, what was the color of the solution without the methylene blue? With the methylene blue? Do these observations shed any additional light on Question 4?
6. What would you predict would happen if the flask were filled to the bottom of the rubber stopper prior to shaking when the experiment was repeated several times in succession?
7. What substances appear to be involved in the reaction?
8. How many steps are involved in the overall process (mechanism of the reaction)? Can you write down these steps? Indicate dextrose (glucose) by G, gluconic acid by GA, methylene blue by MB, colorless methylene blue by MBA, oxygen by O_2, and hydroxide ion by OH^-.

9. Your teacher may provide additional information on the chemistry involved in this series of experiments, to assist you in your understanding of the factors involved in reaction rate studies (chemical kinetics), and the suggested mechanism for this reaction.

10. You should now be able to summarize the factors that affect the rate of a chemical reaction, and explain each of these factors in terms of the kinetic molecular theory of reaction rates.

Implications and Applications

1. Assume that the rates of the reactions involved in cooking double for every 10 °C increase in temperature, and that the three lines on a pressure cooker counter weight correspond to boiling temperatures for water of 110°, 120°, and 130 °C, respectively. On the basis of these facts, complete the following table for the cooking of potatoes in a pressure cooker.

Temperature (°C)	Relative rate of cooking	Cooking time (min)
100	R	80
110		
120		
130		

2. Seashells are used as the raw material for several industrial processes. Consider the reaction between seashells and hydrochloric acid:

$$CaCO_3(s) + 2\,HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Consider the reaction of a single seashell with 1 M HCl(aq). List at least three ways in which you could increase the rate of reaction, and explain on the basis of the kinetic molecular collision theory of reaction rates how each of them increases the rate.

3. Recall that one mole of a substance contains 6.02×10^{23} formula units. For NaOH this would mean 6.02×10^{23} Na^+ ions and 6.02×10^{23} OH^- ions. The molar mass of NaOH is 40.0 g. How would you prepare 5.00 L of 1.5 M NaOH(aq)? How many OH^- ions would there be in (a) 50 mL and (b) 10 mL of this solution?

4. a. What are the sources of error in the mystery flask activity?
   b. Indicate whether the error would lead to high or low values for your measured reaction rates.

5. Did this Mystery Flask activity suggest any additional experiments that would be of interest to perform? How did you and your group members react to studying this reaction system?

Report Sheet
Mystery Flask Experiment

Name: Laboratory Section: Date:

Part I. Initial Observations

Observations on X:

Observations on Y:

Observations on Z:

Conclusions:

Part II. Variation in the Number of Shakes

	Your System (I) Time, s			Other Class Data		
Number of shakes	Initial	Final	Elapsed	II	III	Average Time, s
1						
2						
3						
4						
5						
6						
7						
8						
9						
10						
15						

Draw the plot of number of shakes *vs* time.

Conclusions:

Part III. Variation of the Concentration of Dextrose

Exp	mL Dextrose	mL Water	mL NaOH	Time		
				I	II	III
1	50	0	50			
2	40	10	50			
3	35	15	50			
4	30	20	50			
5	25	25	50			
6	20	30	50			
7	10	40	50			
8	5	45	50			
9	0	50	50			

Plot the number of mL dextrose *vs* time.

Conclusions:

PART IV Variation of the concentration of NaOH

Exp	mL NaOH	mL Water	mL Dextrose	Time		
				I	II	III
1	50	0	50			
2	40	10	50			
3	35	15	50			
4	30	20	50			
5	25	25	50			
6	20	30	50			
7	10	40	50			
8	5	45	50			
9	0	50	50			

Plot the number of mL of NaOH(aq) versus time.

Conclusions:

Part V. Variation of the Concentration of Methylene Blue

Experiment	1	2	3	4	5	6	7
Number of drops of methylene blue							
Time, s							

Plot the number of drops of methylene blue *vs* time.

Conclusions:

Part VI. Variation of Temperature

Experiment	1	2	3	4	5	6	7	8	9	10
Temperature (°C)										
Time, s										

Plot the temperature *vs* time.

Conclusion:

Part VII. Introduction of Selected Gases into the System

Experiment	**1**	**2**	**3**	**4**	**5**	**6**
Gas	O_2	Natural Gas*	Air	N_2	CO_2	He
Time, s						

*The natural gas used in most laboratories is a mixture, which is mostly methane, $CH_4(g)$, or propane, $C_3H_8(g)$.

Conclusions:

LABORATORY ACTIVITY: TEACHER NOTES

Activity 3: The Mystery Flask—An Introduction to Reaction Rates

Major Chemical Concept

This investigation challenges students, working in a cooperative learning environment, to study quantitatively six variables that affect the rate of a fascinating chemical change. This activity can introduce the concept of reaction rates and/or it can be used as an exercise to help students learn how to develop a possible reaction mechanism.

Level

This experiment can be used in part for elementary or middle school science classes. It has been successfully used in all levels of high school and college chemistry classes by selectively choosing the specific experiments and by varying the amount of information provided to the students.

Expected Student Background

Students are expected to be familiar with molarity and the mole concept, have the ability to measure solution volumes with precision, have the ability to make and interpret graphs, and be familiar with kinetic molecular collision theory, and the behavior of gases and solutions

Other Approaches to Using this Laboratory Activity

NOTE: You may wish to carry out this laboratory activity as a small group science project or as a series of student and/or teacher-led demonstrations correlated with extensive discussion. In the past, some teachers have used the initial experiment simply as an experiment in which students collect their individual observations, i.e., *an experiment on observation.*

Time

One or two class periods. Adjust the specific assignments to your available class time blocks.

Safety

1. Demonstrate for the students how to handle a spill of caustic sodium hydroxide on their desks or hands.
2. Show students how to handle gases for carrying out Part G or to generate the needed gases safely.

Materials (For 24 students working in pairs)

Nonconsumables

12 Hot plates (for Part F)
36 400-mL Beakers
24 250-mL Erlenmeyer flasks
24 Stoppers to fit flasks
24 50-mL Graduated cylinders
24 Thermometers, –10 to 110 °C
12 Clocks with sweep second hand, or digital watches which display seconds
12 Burners or hot plates
12 Ringstands and rings
12 Ceramic/wire gauzes

12 3-Foot lengths of rubber tubing with glass portion of medicine dropper attached to one end (*Part VII*)
24 500-mL Reagent bottles with caps (see *Advance Preparation 1*)
Extra flasks and stoppers in reserve for class use

Consumables

1 500-mL 1.5 M NaOH solution
1 500-mL 0.133 M glucose solution
100 mL 0.1% solution of methylene blue (in alcohol) in dropper bottles
12 Ice baths (*Part VI*)
Gas sources (*Part VII*) for oxygen, nitrogen, natural gas, compressed air, carbon dioxide, helium
12 Sheets of graph paper

Advance Preparation

1. Place three 500-mL bottles 1.5 M NaOH, 3 500-mL bottles 0.133 M glucose, and 6 dropping bottles 0.1% methylene blue at each table.

2. Each group will also need a timer with a second hand, two 50-mL graduated cylinders, and a stoppered flask containing about 100-mL Solution X.

3. To make each liter volume of 1.5 M NaOH solution, add 60.0 g NaOH pellets to a 1-L volumetric flask containing about 500-mL distilled water. Stopper the flask and swirl to dissolve. Dilute to one liter. Prepare about 4 L and pour it into large labeled empty bottles.

4. For each liter of 0.133 M dextrose solution, dissolve 24 g dextrose (glucose), $C_6H_{12}O_6$ (molar mass = 180 g/mol) in about 500-mL distilled water in a 1-L volumetric flask. Swirl to dissolve, and dilute to one liter. Again, prepare 4 L of this solution and store in large labeled empty bottles.

5. Alternatively, you can prepare 6 M NaOH (240 g/L) and 0.532 M dextrose (glucose; 96 g/L), and then dilute each 1:3 with distilled water.

6. To prepare 0.1% methylene blue solution, dissolve 10 g of the solid in 100 mL ethanol (95% by mass), and transfer to dropper bottles.

7. Solutions X, Y, and Z are made by premixing equal volumes (50 mL each) 1.5 M NaOH and 0.133 M dextrose solutions. Solutions Y and Z can be prepared 1-5 days in advance and then subdivided into labeled flasks for each group. The same Y and Z solutions can be used for all classes. Note that Y and Z must be prepared at least one day prior to the laboratory activity.

   Solution X: Just prior to each class, mix equal volumes of 1.5 M NaOH and 0.133 M dextrose (glucose) solutions in a large stoppered container. For 12 laboratory groups, approximately 600 mL each will be necessary. Then add sufficient drops of 0.1% methylene blue solution (30-plus drops) so that, on shaking, an easily observed blue color forms. Then pour 100-mL portions of this mixture into 250-mL flasks labeled X. This procedure assures that all groups will have the same original solution.

   Solution Y: Repeat the same procedure for preparing Solution X, except do this one or more days before the day of the activity. Let this solution stand at least overnight.

   Solution Z: Repeat the procedure for Solution Y, omitting the methylene blue solution. Again, let stand at least overnight.

8. You may wish to demonstrate the experiment in which the flask is filled completely with the solution system and several sequences of shakes are performed. (The air [O_2] concentrations will soon drop to zero.)
9. Student groups or partners should be assigned one or more variables to study. The class data can be recorded on a master list, and copies can be run off for each student or the data can be written down on the blackboard, poster board or manila folder. Each student should collect a complete set of data for their report sheets.
10. A post-laboratory class discussion is an excellent medium for both understanding the specific reaction and the general factors governing reaction rates.

References

Campbell, J. A. (1963). *Journal of Chemical Ed*ucation, *40*, 578.

Shakhashiri, B. Z. (1985). *Chemical Demonstrations*, Volume 2, *6*(9).

11. Possible sources for gases for *Part VII*:

Lecture bottles
Regular cylinders (tanks)
Large heavy balloons filled from local college or industry
Laboratory gas supply
Medical oxygen supply
Butane and/or carbon dioxide cartridges
Supplier of balloons (helium)
Carbon dioxide fire extinguisher or dry ice

Pre-Laboratory Discussion

Discuss the term *rate* by discussing and/or demonstrating one or more of the analogies listed later in this module. It might be appropriate to demonstrate a reaction other than the one used in this activity as a focal point for discussing reaction rates. *Part I* of this laboratory activity could be done very early in the course as an exercise in careful observation, similar to the goal of the CHEM Study candle activity.

Teacher-Student Interaction

Move around the laboratory throughout the period. Watch for the use of proper technique in using the equipment, observation of safety precautions, participation of all students in the group, and how specific teams function together. Ask questions to check student understanding to encourage further exploration and to facilitate maximum learning and motivation. Check to see that students properly record data with the correct number of significant figures. You may find that the first piece of data in a series of runs may be inconsistent with the rest of the data.

Anticipated Student Results

Part A

Upon shaking, the X solution turns blue. The blue color gradually decolorizes until a ring of blue or swirls of blue remain near the surface. Hopefully, students will eventually deduce that the shaking brings oxygen into the solution, which oxidizes the colorless form of methylene blue to the blue form.

Both Solutions Y and Z are usually a golden brown color indicating that, given enough time, both the catalyzed reaction system Y and the uncatalyzed system Z will eventually react to the same degree as they achieve equilibrium.

Part B. Number of Shakes

At a temperature of 20 °C, the following data were obtained:

Shakes	1	2	3	4	5	6	7	8	9	10
Time, s	12.5	17.8	23.3	27.0	33.5	57.0	58.0	60.0	67.0	68.0

Clearly there is a maximum amount of air that can be dissolved by shaking. The type of shaking is also a variable, but each group should have self-consistent results. The graph of number of shakes *vs.* time should level off.

Part C. Concentration of Glucose (Dextrose)

Number of shakes	3
Temperature	25 °C
Drops of methylene blue	7

mL Dextrose	**mL H_2O**	**mL NaOH**	**Time, s**
35	15	50	19.8
30	20	50	43.0
20	30	50	77.0
10	40	50	121

Part D. Concentration of NaOH

Number of shakes	3
Temperature	23 °C
Drops of methylene blue	7

mL NaOH	**mL H_2O**	**mL Dextrose**	**Time, s**
50	15	35	19.8
50	20	30	43.0
50	30	20	77.0
50	40	10	121

Part E. Concentration of Methylene Blue

mL NaOH	50
mL Glucose	50
Temperature	24 °C

Drops of Methylene Blue

Shakes	**2**	**3**	**4**	**7**
1	10.1 s	—	—	—
2	25.7 s	15.3 s	—	—
3	38.7 s	20.9 s	25.9 s	13.3 s
5	46.9 s	39.3 s	33.4 s	27.2 s
7	62.3 s	45.7 s	40.0 s	34.4 s
10	74.6 s	54.4 s	42.0 s	40.4 s
15	74.8 s	67.8 s	57.0 s	52.1 s
20	78.9 s	75.2 s	58.7 s	56.2 s

Part F. Temperature

mL NaOH	50
mL Dextrose	50
Number of shakes	2
Drops of methylene blue	4

Temperature (°C)	Time, s
10 (283 K)	216
25 (293 K)	27
40 (313 K)	3

Add methylene blue after the solution mixtures have equilibrated to the desired temperatures and then shake. Maintain the flasks at the desired temperatures after shaking by returning the flasks to the temperature baths. A complication is the decreasing solubility of air and O_2 with temperature. The number of shakes can be varied along with the temperature.

Part G. Dissolved Gases

mL NaOH	50
mL Dextrose	50
Number of shakes	2
Drops of methylene blue	7
Temperature	23 °C

Gas	Time, s
O_2	150
CH_4	No color change upon shaking
CO_2	No color change upon shaking

Notice the much longer time with pure oxygen. The other two gases displace the oxygen from the solution; thus oxidation of the colorless form of methylene blue cannot occur.

Determining the mechanism: The overall mechanism in general form is shown in Figure 23.

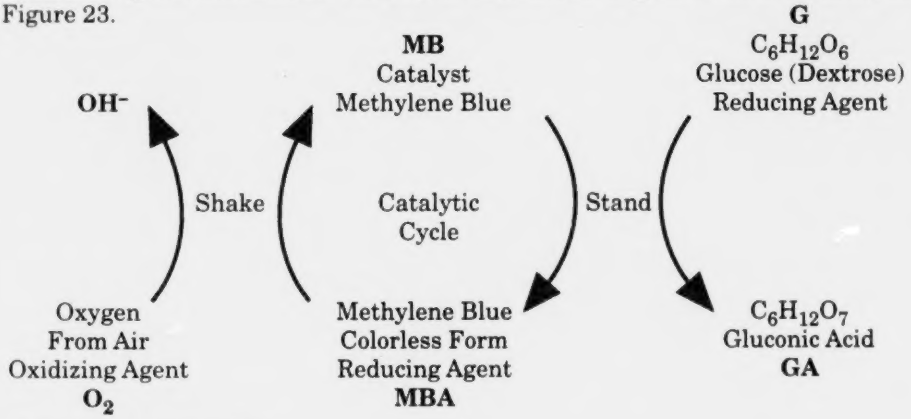

G
$C_6H_{12}O_6$
Glucose (Dextrose)
Reducing Agent
MB
Catalyst
Methylene Blue
OH⁻
Shake
Catalytic
Cycle
Stand
Oxygen
From Air
Oxidizing Agent
O_2
Methylene Blue
Colorless Form
Reducing Agent
MBA
$C_6H_{12}O_7$
Gluconic Acid
GA

Figure 23. Mechanism.

A possible simplified mechanism is:

Step 1	$MB + G + OH^-$	$\rightarrow$	$MBH + GA$	Slow
Step 2	$^1/_2O_2(g)$	$\rightarrow$	$^1/_2O_2(aq)$	Moderate
Step 3	$MBH + {}^1/_2O_2(aq)$	$\rightarrow$	$MB + OH^-$	Fast
Overall	$G + {}^1/_2O_2(g)$	$\rightarrow$	GA	

Notice that both MB and OH^- are functioning as catalysts (reactants in Step 1 and products in Step 2). Clearly this proposed mechanism needs further investigation and thought since Step 1 is outlined as a termolecular reaction.

Answers to Implications and Applications

1.

Temperature (°C)	Relative rate of cooking	Cooking time (min)
100	R	80
110	2R	40
120	4R	20
130	8R	10

2. a. Increase the concentration of HCl (to 3 M, 6 M or 12 M, for example).

   b. Grind up the seashell, which increases the surface area and facilitates more collisions of the H_3O^+ ion with the $CaCO_3$.

   $$CO_3^{2-}(aq) \rightarrow HCO_3^-(aq) \rightarrow H_2CO_3(aq) \rightarrow CO_2(g) + H_2O(l)$$

   c. Increasing the temperature will increase both the number of collisions and the average collisional energy.

3. $$\text{Molarity} = \frac{\text{Moles NaOH}}{\text{Liters of solution}} = \mathbf{1.5\ M\ NaOH}$$

$$5.00\ \text{L} \times \frac{1.5\ \text{mol NaOH}}{1\ \text{L solution}} \times \frac{40.0\ \text{g NaOH}}{1\ \text{mol NaOH}} = \mathbf{3.00 \times 10^2\ g\ NaOH}$$

a. $$50\ \text{mL} \times \frac{1\text{L}}{1000\ \text{mL}} \times \frac{1.5\ \text{mol NaOH}}{1\ \text{L NaOH}} \times \frac{6.02 \times 10^{23}\ OH^-}{1\ \text{mol NaOH}} = \mathbf{4.5 \times 10^{22}\ OH^-\ ions}$$

b. $$10\ \text{mL} \times \frac{1\ \text{L}}{1000\ \text{mL}} \times \frac{1.5\ \text{mol NaOH}}{1\ \text{L NaOH}} \times \frac{6.02 \times 10^{23}\ OH^-}{1\ \text{mol NaOH}} = \mathbf{9.0 \times 10^{21}\ OH^-\ ions}$$

4.i.a. Measuring the volumes of solutions with graduated cylinders. It is a low precision measuring device in terms of reading volumes, and the fact that solution remains in the cylinder after pouring.

b. With the careful use of the same graduate, the error will be 2-4%, high or low.

ii.a. The amount of dissolved oxygen brought into the solution will vary with the type of shake.

b. The relative errors from trial-to-trial can be either plus or minus.

iii.a. The time measurement will generally have a high precision, and thus be a negligible error.

iv.a. Since reaction rates roughly double for each 10 °C increase in temperature, errors in reading and/or maintaining a constant solution temperature will lead to significant errors.

b. Either positive or negative.

v.a. The added drops of methylene blue may vary in size;

b. Leading to positive or negative errors.

vi. In the study of the number of shakes (*Part II*) the rubber stopper should be removed between successive runs to provide a relatively constant concentration of O_2 in the gas phase.

vii. It is best not to carry out the experiment in sunlight or under fluorescent lights since the dye is photochemically changed to a violet color.

5. See *Extensions* below.

Extensions

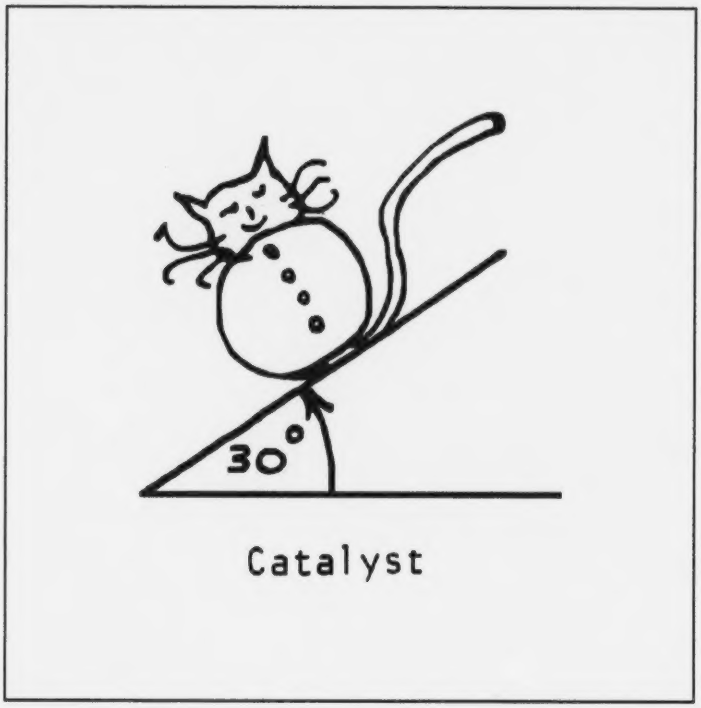
1. Try other reducing sugars, such as mannose, galactose, lactose, maltose or ribose.
2. Expose the reaction system to strong light on an overhead projector. A photochemical change occurs. The light photons will transform irreversibly the blue form of methylene blue to a violet compound.
3. Explore other catalysts or combinations of catalysts, which will give other single or multiple color changes, *e.g.*, indigo carmine or rhodamine B.
4. Explore ozone, O_3, or dinitrogen monoxide, N_2O, as the dissolved gas. Ozone, a stronger oxidizing agent than oxygen may act similarly to oxygen or it might completely destroy the dye. Like oxygen, N_2O supports combustion and may also convert the colorless (leuco) form of methylene blue to the blue form. *CAUTION! in using these gases.*
5. Carry out the reaction ranging the size of the Erlenmeyer flasks from 125 mL, 250-mL, 50-mL, 1000-mL, 2000-mL. (O_2 moles increase with increasing size of flask.)
6. You may wish to have students investigate the following *Purple Bottle* activity.

"Purple Bottle" Laboratory Activity

Add 100 mL 2-propanol to a flat-bottom 250-mL flask and add 0.100 g of benzoin. Stopper and shake to form a solution. Add a solution made by dissolving 2 g KOH in 5 mL water. Stopper, shake and let stand. The solution will turn a grape color. Alternately, remove the stopper, shake to bring oxygen into the system and then let the system stand. Additional increments of solid benzoin may be added to continue the reaction. Note the yellow color of the final product in the solution. The purple (or grape) color is due to a complex of benzil, a diketone, which is a product of auto-oxidation of benzoin.

Post-Laboratory Activities

1. Have each group of students present its generalizations from the data, and explanation of the data in terms of reaction rates and the collision model of reaction rates.
2. Go over graphs of the data.
3. Assist the class in the development of an overall explanation of the reaction steps (mechanism). Attempt to lead students with suggestions rather than yielding to the temptation to take over and explain all the data. The demonstration books by Summerlin *et al.* and Shakhashiri would be helpful in both interpreting the results and in suggesting experimental extensions.
4. Review the steps in the experimental procedures in terms of sources of error.


30°
Catalyst

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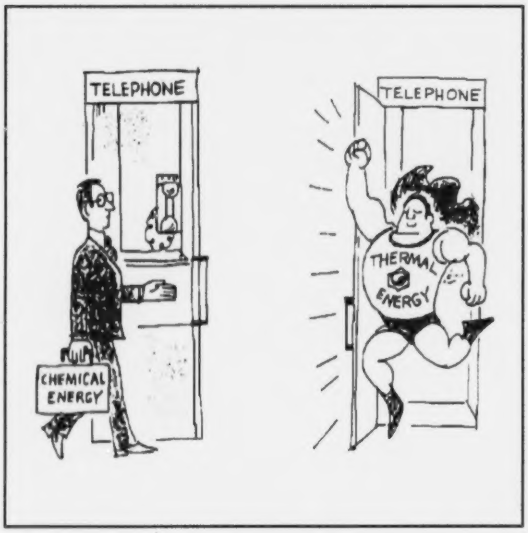
TELEPHONE
CHEMICAL ENERGY
TELEPHONE
THERMAL ENERGY

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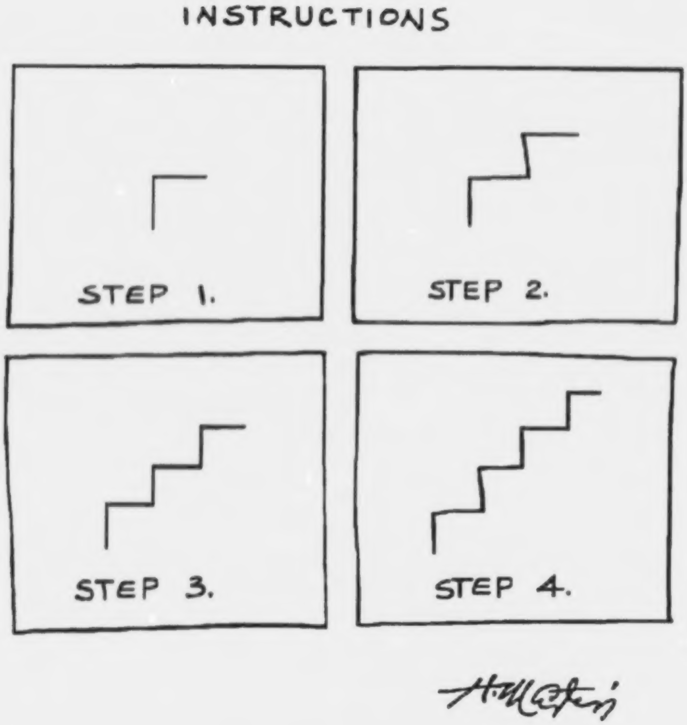
INSTRUCTIONS
STEP 1.
STEP 2.
STEP 3.
STEP 4.

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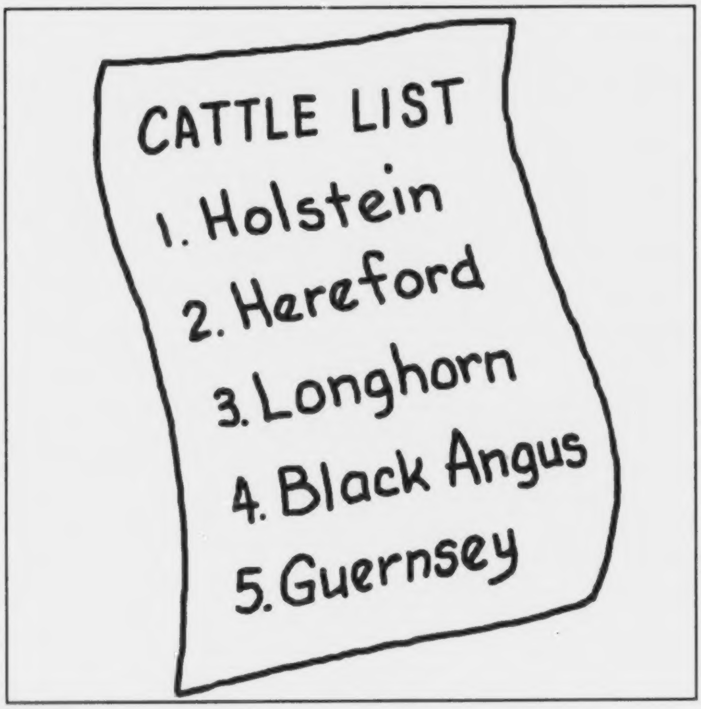

CATTLE LIST

1. Holstein
2. Hereford
3. Longhorn
4. Black Angus
5. Guernsey

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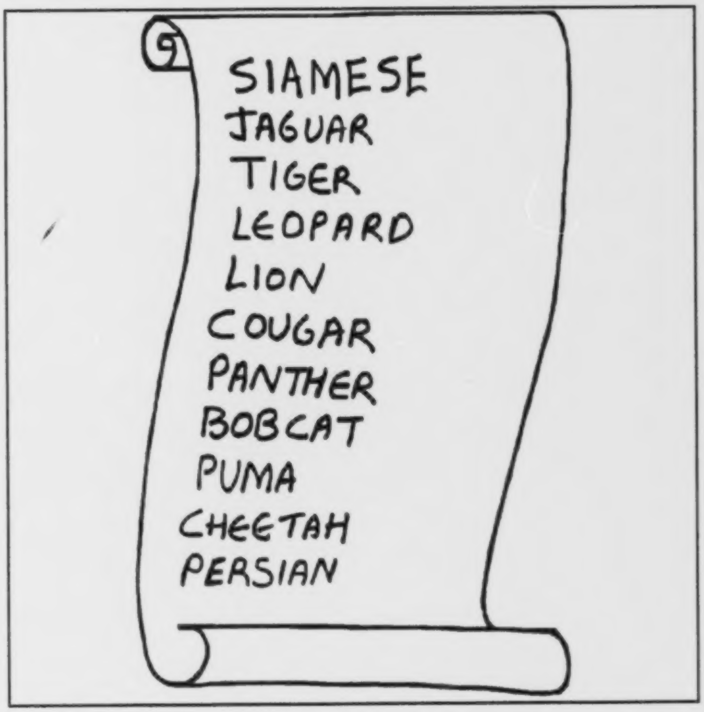
SIAMESE
JAGUAR
TIGER
LEOPARD
LION
COUGAR
PANTHER
BOBCAT
PUMA
CHEETAH
PERSIAN

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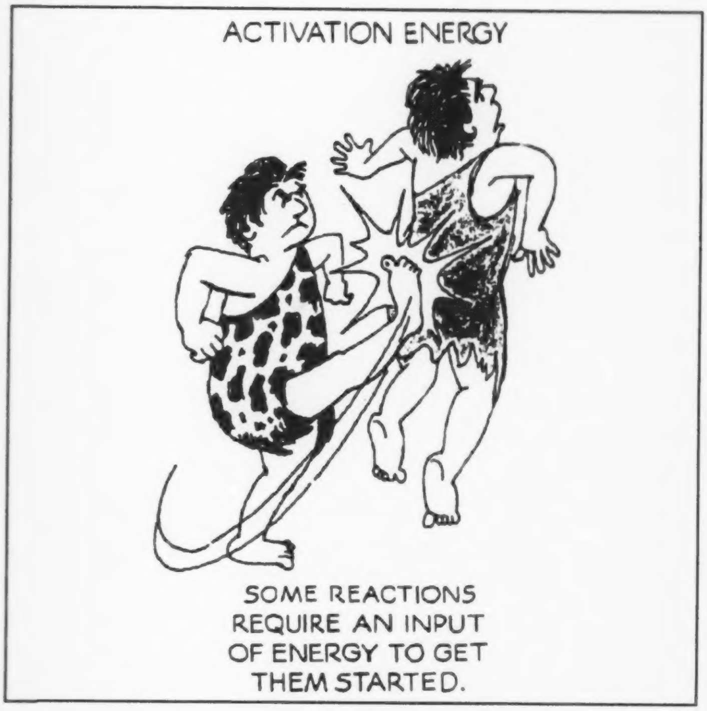
ACTIVATION ENERGY
SOME REACTIONS REQUIRE AN INPUT OF ENERGY TO GET THEM STARTED.

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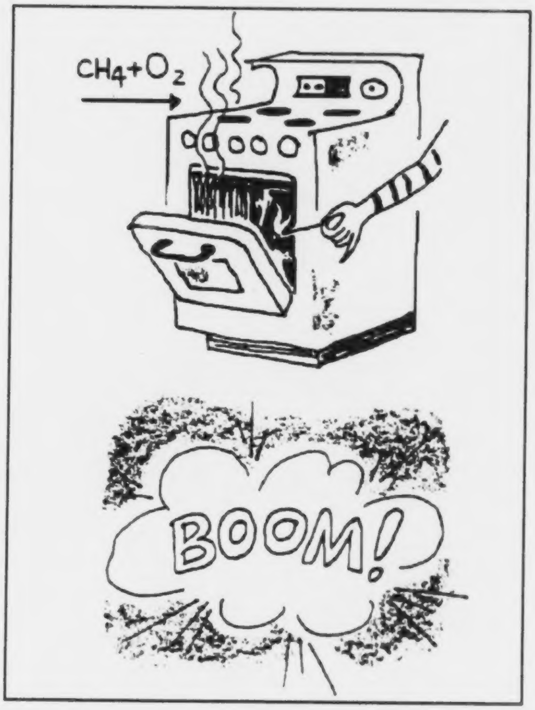

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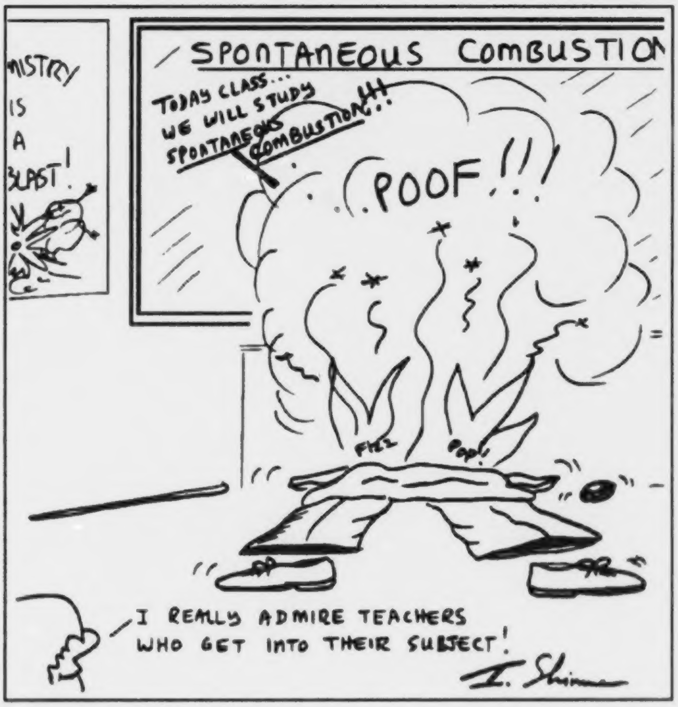
MISTRY IS A BLAST!
SPONTANEOUS COMBUSTION
TODAY CLASS... WE WILL STUDY SPONTANEOUS COMBUSTION!!!
POOF!!!
FIZZ
POP!
I REALLY ADMIRE TEACHERS WHO GET INTO THEIR SUBJECT!

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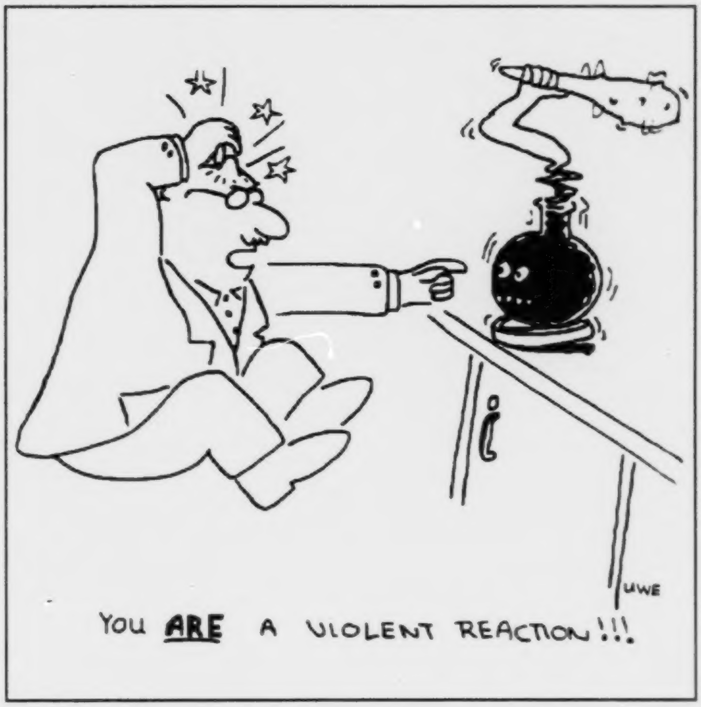
UWE
YOU ARE A VIOLENT REACTION!!!

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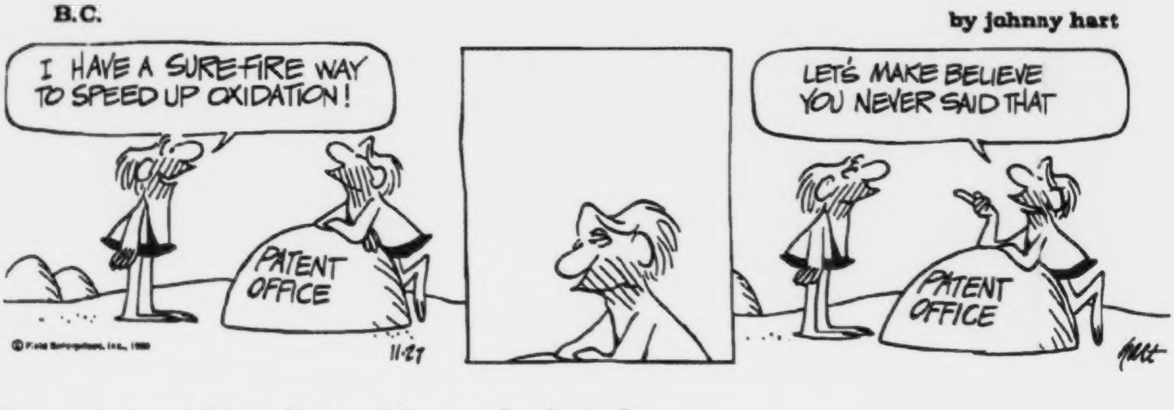
B.C.
by johnny hart
I HAVE A SURE-FIRE WAY TO SPEED UP OXIDATION!
PATENT OFFICE
11-27
LET'S MAKE BELIEVE YOU NEVER SAID THAT
PATENT OFFICE

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A SourceBook Module

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THE CHEMISTRY OF ROCKS, MINERALS, AND GEMS

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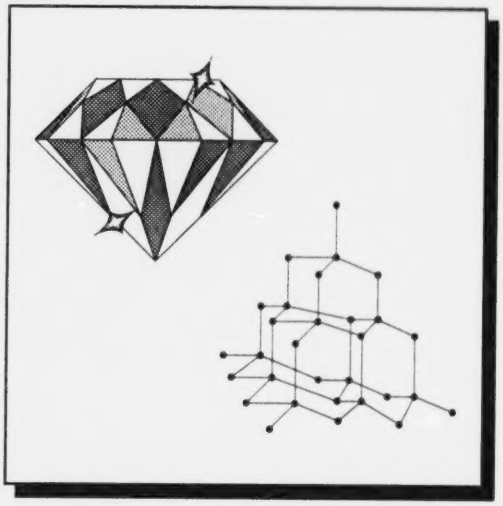


ChemSource

Instructional Resources for Preservice and
Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

The planet earth is composed of three distinct regions, the core, the mantle, and the crust (Figure 1). The core is believed to be mainly a sphere of iron and nickel which extends about 3500 km from the center. There is an inner core that appears to be solid, and an outer core that is molten. The mantle is thought to be mostly the silicate mineral olivine, $(Fe,Mg)_2SiO_4$. This extends another 3000 km. The region about which we know the most, the crust, has a thickness of 5 to 100 km. It is composed of the hydrosphere and the lithosphere. The lithosphere, which is hard, brittle, and relatively cool, rides on the upper portion of the mantle, the asthenosphere, which is hot, plastic, and semiliquid. Heat and pressure only partially explain the existence of the semiliquid layer between two solid layers. A jelly sandwich is an effective analogy.

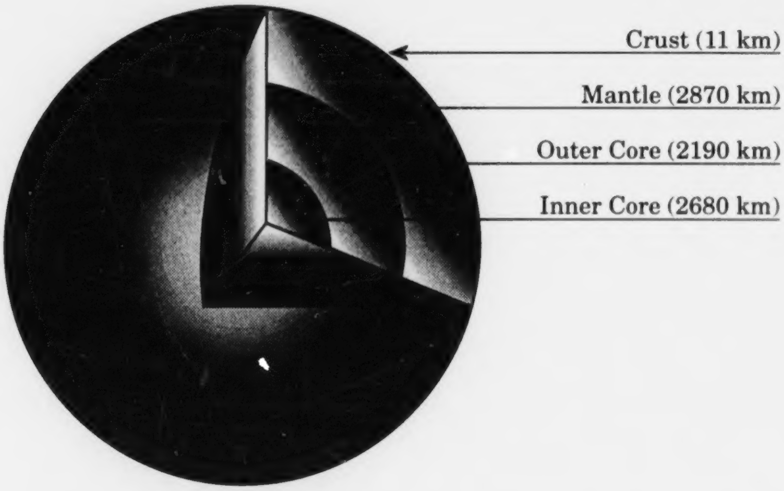


Figure 1. Regions of the earth.

Figure 2 shows how the temperature of the regions varies. The temperature increases from the crust to the inner core.

Depth, km	Temperature, °C
30	500
100	1100
200	1400
1000	2000
3000	2700
6000	3000

Figure 2. Temperatures beneath the earth's surface.

The ten most abundant elements in the crust are shown in Figure 3. Oxygen is the most abundant nonmetal and aluminum is the most abundant metal.

Oxygen, the most abundant element, comprises about 20% of the atmosphere, but most of the oxygen in the crust is found in the lithosphere not as an element, but combined with other elements. The second most abundant element, silicon, is found combined with oxygen in silicon dioxide and over a thousand different silicates. The numerous variations are due to the ability of silicon to bond to oxygen to form SiO_4^{4-} tetrahedra and the ability of these tetrahedra to join by sharing a common oxygen.

Element	**Mass %**	**Mole %**
Oxygen	49.4	42.5
Silicon	25.8	25.4
Aluminum	7.5	7.7
Iron	4.7	2.2
Calcium	3.4	2.5
Sodium	2.6	3.0
Potassium	2.4	1.7
Magnesium	1.9	2.2
Hydrogen	0.9	12.4
Titanium	0.6	0.2

Figure 3. Most abundant elements in the crust.

Silicon's position in the mineral kingdom is analogous to that of carbon in the animal and vegetable kingdoms. Silicon, like carbon, is tetravalent, and forms various giant molecules and polymers. Unlike carbon, however, silicon cannot form endless bonds with its own atoms and has limited ability to form multiple bonds. The strikingly different properties of $SiO_2(s)$ and $CO_2(g)$ illustrate this dissimilarity. Silicon dioxide is a network of joined tetrahedral SiO_4 units, while carbon dioxide consists of discrete molecules, O=C=O. The silicates also contain the SiO_4 units joined to form chains, layers, sheets, or three-dimensional arrangements. Synthetic silicone polymers also contain Si–O bonds, as well as carbon-containing organic groups.

The uses of silicon compounds are related to their thermal stability and inertness. The bond energy of the Si–O bond is 430 kJ/mol compared to 190 kJ/mol for Si–Si and 340 kJ/mol for C–C. Although many elements are found as silicates, silicates are not generally useful as sources of these elements because of the difficulty of extracting them. Silicates are primarily used with only small modifications of their structure—as glasses, ceramics, and cement. Because of the strength of the Si–O bond, synthetic silicones—$(R_2SiO)_n$ with R = CH_3, C_2H_5, C_6H_5—are used as lubricants, waxes, greases, and water repellents.

Minerals are inorganic solids of definite composition found in the crust. They include the silicate minerals as well as uncombined elements (*e.g.*, Cu, Ag, Au, and S) and compounds such as oxides, carbonates, sulfates, sulfides, halides, and phosphates. Iron, the fourth most abundant element in the earth's crust, is of great economic importance as a structural material. The major sources of iron are ores containing 30–40% iron in the form of hematite, Fe_2O_3, or magnetite, Fe_3O_4. Aluminum is also of considerable economic importance. It is chiefly obtained from bauxite, a mixture of alumina, Al_2O_3, with major impurities of Fe_2O_3 and SiO_2. After separation from these impurities the alumina is reduced by electrochemical means. Calcium, the fifth most abundant element is found as limestone, marble, and chalk (all varieties of $CaCO_3$), anhydrite, $CaSO_4$, gypsum, $CaSO_4 \cdot 2H_2O$, fluorite, CaF_2, and apatite, $Ca(F,Cl)_2 \cdot 3Ca_3(PO_4)_2$. Some of these materials (limestone, marble, and gypsum) are used without modification for structural purposes. Modifications and combinations with silicates produce plaster of Paris, cement, ceramics, and bricks.

Although the earth seems permanent to us, it is constantly changing. Movement of the solid lithosphere on the plastic asthenosphere causes cataclysmic changes such as earthquakes and volcanoes. The exposure of the rocks at the surface to atmospheric gases, water, and temperature changes also brings about significant alteration of the crust. These changes are called weathering and ultimately result in converting rocks to soil. The geologist divides weathering into mechanical and chemical. Some chemists might argue that even mechanical weathering is primarily chemical since most of the mechanical changes are brought about by the actions of water. Freezing water causes rocks to break and chip. Water dissolves minerals and transports them to distant locations and deposits them. Mineral ore deposits are created when water concentrates minerals in a small region. Chemical changes result when water, oxygen, and carbon dioxide react with rocks, producing new compounds and often causing rocks to crumble. Carbon dioxide dissolves in water, producing an acidic solution that reacts with some minerals. These processes have been active for millions of years and result in a land surface covered with rock fragments.

PLACE IN THE CURRICULUM

If this module is taught as a separate unit, the student should have some background in equilibrium, chemical bonding, solutions, and molecular geometry. You may wish to use one or more of these activities to apply the chemistry in other modules.

CENTRAL CONCEPTS

1. Minerals can be formed by crystallizing from a melt or by deposition from water.
2. Minerals can be identified by chemical tests such as flame tests, bead tests, reactions with acids, and fluorescence.
3. Bonding determines structure, and structure determines properties.
4. Rocks are mixtures of minerals with variable composition.
5. Organic, inorganic, and physical processes called weathering result in changing rocks to soil.
6. Gems are often minerals that contain small amounts of transition elements as impurities.

RELATED CONCEPTS

1. Covalent network solids have atoms covalently bonded in three-, two-, or one-dimensional networks (see *Chemical Bonding* module).
2. Chemical reactions can reach a state of dynamic balance: equilibrium (see *Equilibrium* module).
3. LeChatelier's Principle: If a system is stressed it will tend to react in a way to counteract the stress (see *Equilibrium* module).
4. The solubility of a precipitate can be altered by adding a common ion or adjusting the pH (see *Solubility and Precipitation* module).
5. Accumulated data allow prediction of the degree of solubility of many substances (see *Solubility and Precipitation* module).
6. Properties of network solids can be related to their structures (see *Materials Science* module).
7. Concepts of solution concentration (concentrated, dilute, supersaturated, saturated, unsaturated, and solubility curves) are reviewed in *Solubility and Precipitation* module.

RELATED SKILLS

1. Predict bond type from a knowledge of the Periodic Table and electron configurations.
2. Understand the three-dimensional character of network solids.
3. Analyze data from qualitative analysis activities.
4. Translate from two-dimensional drawings to three-dimensional models.

PERFORMANCE OBJECTIVES

After completing their study of rocks, minerals and gems, students should be able to:

1. define a mineral.
2. distinguish between the two main mineral groups.
3. identify the elements found most abundantly in common minerals.
4. prepare a crystal.
5. distinguish between mineral and gemstone.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Making Crystals

Introduction

Atoms, molecules, and ions can fit together in a regular pattern to form a crystal. There are several types of crystals and crystalline structures. One of the ways to grow crystals is to allow water to evaporate so that the ions or molecules can come together to form crystals.

Purpose

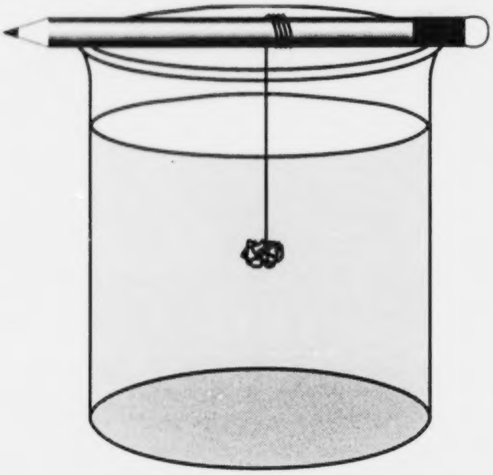
To investigate sugar and salt crystals. To produce salt crystals by allowing sodium ions and chloride ions to come together in solution. To produce sucrose crystals (sometimes called rock candy) by allowing sugar molecules to come together in solution.

Safety

Wear protective goggles throughout the laboratory activity.

Procedure

Part I. Preparing Salt Crystals



1. Fill a 250-mL beaker about three-fourths full of water.
2. Add salt (about 20 g) to the water until no more will dissolve.
3. Heat the solution (to about 90 °C), but do not boil it.
4. While the solution is hot, add salt (about 15 g) until no more will dissolve. Remove the beaker from the hot plate. Be sure that no undissolved salt remains in the bottom of the beaker.
5. Suspend a string in the solution by tying the string onto a pencil, soda straw or wood splint and suspending it across the top of the beaker.
6. Cover the beaker with a watch glass but do not seal it. Put the beaker where it will not be disturbed for a few days. Make observations daily. Record your observations.

Part II. Preparing Sugar Crystals

1. Boil 150 mL water in another 250-mL beaker.
2. Add as much sugar (about 110 g) as will dissolve in the hot water.
3. Repeat Steps 5 through 7 from Part A.
4. Compare both types of crystals. Look at them with a microscope if available.
5. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. Describe the results of this activity. What did the crystals look like? Draw diagrams to show the crystals.
2. What are the characteristics of each type crystal?
3. What purpose does the string serve?

Implications and Applications

1. Imagine how the particles of each substance must be arranged to produce the crystals you described. Draw a "particle nature of matter" picture to describe what you imagine.
2. If sugar crystals are made of sugar molecules, and salt crystals are made up of sodium ions and chloride ions, how would the particle pictures differ?
3. Explain how the process of crystallization might have been involved in forming deposits of minerals.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Making Crystals

Major Chemical Concept

Substances crystallize in predictable, regular patterns.

Level

General or Basic

Expected Student Background

None.

Time

45 min the first day and 5 min on successive days (for observations).

Safety

Students should be reminded not to taste anything in the laboratory, even things that can ordinarily be ingested.

Materials (For 24 students working in pairs)

Nonconsumables

24 Beakers, 250-mL
12 Hot plates

Consumables

Sugar, 600 g
Salt, 480 g
12 Kite strings, 10 cm

Advance Preparation

No advance preparation is needed. *NOTE*: A heavy-gauge pre-soaked string works best since it will tend to sink.

Pre-Laboratory Discussion

This activity might be preceded by showing shapes of some naturally occurring crystals and speculating about what might happen when solids are crystallized by evaporation. A discussion of the changes in concentration of a solute that occur when the water evaporates will help students understand what happens in the activity. The terms "dissolve" and "saturated" should be reviewed.

Teacher-Student Interaction

This activity is easy to perform and little interaction is needed. If necessary, students can be reminded that if solid remains in the beaker after heating, the solution is saturated, and no more solute should be added.

Anticipated Student Results

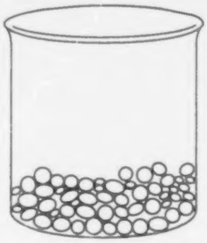
Both sugar and salt crystallize in the cubic form. Students may observe several small crystals or fewer larger ones.

Answers to Data Analysis

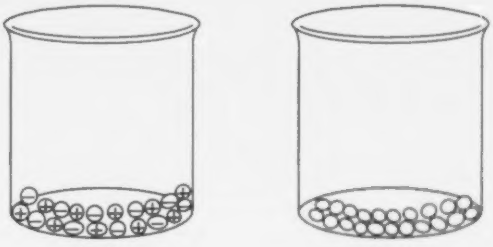
1. Small cubes.
2. The particle at the lattice points for sugar is a molecule of sugar; in the sodium chloride crystal the particles are sodium ions and chloride ions.
3. To suspend the crystal so that it has a greater opportunity to serve as a growth site, and to minimize crystal distortion by resting in the bottom of the beaker.

Answers to Implications and Applications

1.



2.



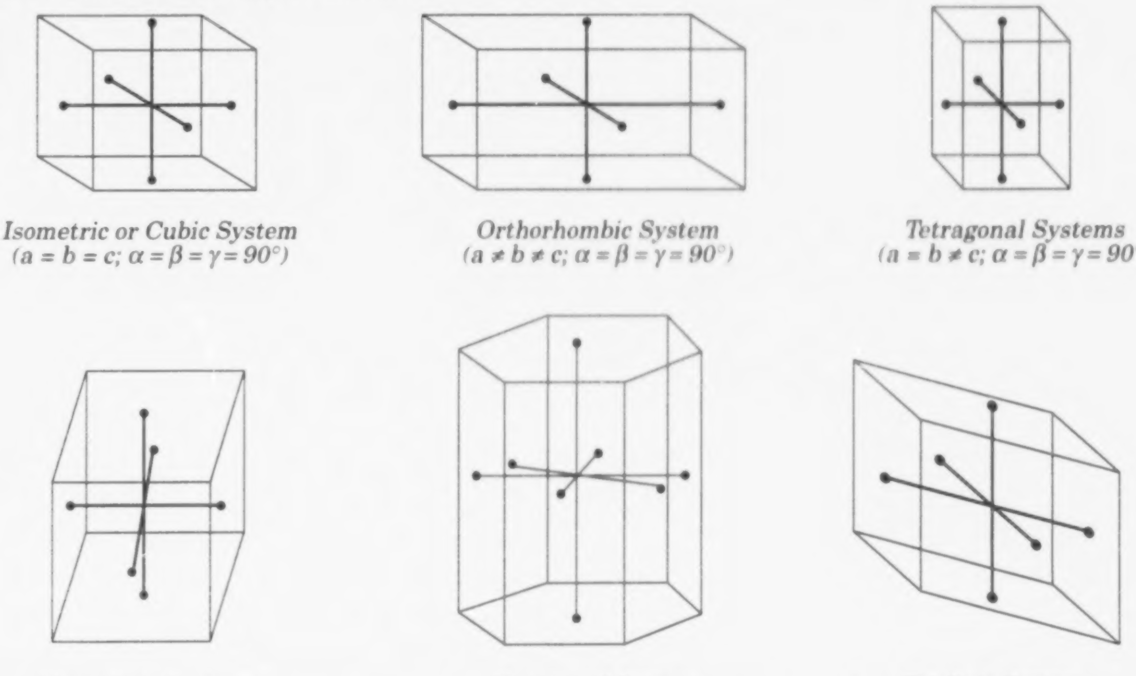
Salt

Sugar

3. As water evaporates, the solution becomes saturated and crystals begin to form.

Post-Laboratory Activities

After the crystals begin to grow, a discussion of equilibrium would be appropriate if this topic was previously introduced. The six crystal systems (Figure 4) and examples from the mineral world may be introduced.



Isometric or Cubic System
(a = b = c; α = β = γ = 90°)

Orthorhombic System
(a ≠ b ≠ c; α = β = γ = 90°)

Tetragonal Systems
(a = b ≠ c; α = β = γ = 90°)

Monoclinic Systems
(a ≠ b ≠ c; α = β = 90° ≠ γ)

Hexagonal System
(a = b ≠ c; α = β = 90°; γ = 120°)

Triclinic System
(a ≠ b ≠ c; α ≠ β ≠ γ ≠ 90°)

Figure 4. Six basic crystal systems: isometric or cubic, orthorhombic, tetragonal, monoclinic, hexagonal, triclinic; a, b, c = edge lengths;** α, β, γ **= internal angles.

Extensions

1. Students may wish to continue crystal growing with other types of compounds, *e.g.*, copper sulfate, alum, and Epsom salts. If students wish to keep the crystals they make, cover them with clear nail polish to prevent reaction with the atmosphere.
2. Have students compare the basic crystal systems and show how they differ from one another.

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: Identification of Minerals by Borax Bead Tests and Flame Tests

Introduction

Two simple tests that can be used to identify the metal ion in minerals are flame tests and borax bead tests. Both use only simple equipment and require a short amount of time. Borax beads can be made in different colors depending on whether a reducing or an oxidizing flame is used.

Purpose

To identify metal ions through flame tests and borax bead tests.

Safety

Wear protective goggles throughout the laboratory activity.

Procedure

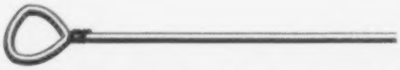
Prepare a data table to record your results.

Part 1. Flame Tests

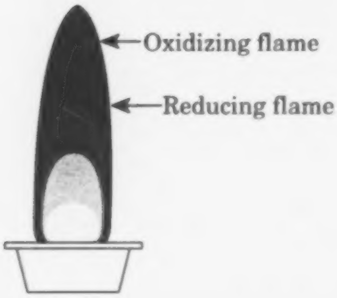
1. Place pea-sized amounts of the salts KCl, NaCl, LiCl, and $CuCl_2$ in separate watch glasses.
2. Adjust the burner so that a distinct blue cone is present.
3. Clean a platinum wire by dipping it into 6 M HCl and then holding it in the flame until no color is produced.
4. Dip the platinum wire in the KCl, and hold it in the flame. Record the color.
5. Clean the wire as before, and repeat the tests with the other salts.

Part II. Borax Bead Tests

1. Twist the end of the platinum wire to form a loop as shown.



2. Place about 1 gram of $CuCl_2$, $CoCl_2$, $Fe(NO_3)_3$ MnO_2, and borax in separate watch glasses.
3. Heat wire first in flame and then dip the loop of the wire in the borax, picking up a small amount. Heat it until a glassy bead is formed.



4. Dip the bead in the $CuCl_2$ and return to the flame's outer core (oxidizing portion). This portion of the flame is inside the luminous part of the flame about a centimeter above the inner blue cone. Record the color, and repeat the process in the reducing portion of the flame. This region is at the tip of the inner blue cone.
5. Crush the bead from the platinum wire, and clean the wire with 6 M HCl as described above.
6. Bring a magnet close to the bead. Observe for attraction.

7. Repeat the bead test with the remaining solids.
8. Obtain an unknown from your teacher. Perform flame tests and bead tests.
9. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

Use your data table to determine the identity of your unknown.

Implications and Applications

Can a mineral be identified by flame and bead tests alone? Why or why not?

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: Identification of Minerals by Borax Bead Tests and Flame Tests

Major Chemical Concept

Metal ions can be identified by flame and borax bead tests.

Expected Student Background

Students should be able to:

1. use a burner safely.
2. make a data table.
3. interpret data in a table.

Time

45 min

Safety

Remind students about hazards of operating burners and using 6 M HCl.

Materials (For 24 students working in pairs)

Nonconsumables

12 Burners and tubing
12 Platinum wires in glass or nichrome wires attached to balsa wood
60 Watch glasses
12 Test tubes, 13- x 100-mm
12 Stirring rods or spatulas (for crushing the borax bead)

Consumables

Sodium chloride, $NaCl$, 1.2 g
Potassium chloride, KCl, 1.2 g
Lithium chloride, $LiCl$, 1.2 g
Copper(II) chloride, $CuCl_2$, 1.2 g
Cobalt(II) chloride, $CoCl_2$, 1.2 g
Manganese oxide, MnO_2, 1.2 g
Iron(III) nitrate, $Fe(NO_3)_3$, 1.2 g
Borax, 6.0 g
6 M Hydrochloric acid, HCl, 150 mL (76 mL conc. HCl diluted to 150 mL)
Cobalt glass *[OPTIONAL: This glass is sometimes helpful in observing the potassium flame test by elminating the interference of the sodium orange-yellow flame.]*

Advance Preparation

Prepare 6 M hydrochloric acid. Be sure to add acid to water.

As the wire deteriorates with use, cut off a portion to get a fresh sample.

Prepare unknowns from samples used.

Pre-Laboratory Discussion

Remind students of safety hazards. Demonstrate the oxidizing and reducing portions of the flame.

Teacher-Student Interaction

Students sometimes need to be reminded to make observations as quickly as possible after the wire has been put into the flame. When the wire becomes red hot, the flame will appear orange regardless of the substance being tested.

Anticipated Student Results

Flame Tests	
KCl	Lavender (difficult to see)
NaCl	Yellow-orange
LiCl	Red
$CuCl_2$	Green

Borax Bead Tests		
	Oxidizing color	**Reducing color**
$CuCl_2$	Green	Red
$CoCl_2$	Blue	Blue
$Fe(NO_3)_3$	Yellow	Green
MnO_2	Red	White

Answers to Data Analysis

Will depend on identity of student's unknown.

Answers to Implications and Applications

These tests identify only the metal ion. Other tests such as density, color, cleavage, and scratch test will give further information to help in identification (see MECC computer simulation in *Media*).

Post-Laboratory Discussion

Relate these tests to similar tests that might be made on minerals.

Extension

If time permits, obtain mineral samples from the geology or earth science teacher and allow students to perform the tests with them. Check any minerals for possible hazards.

Assessing Laboratory Learning

The extension may be used for assessment. If you do not wish to use actual minerals, prepare unknowns from the salts used in the activity.

DEMONSTRATIONS

CAUTION: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Growing Crystals in Gels

Description

In nature minerals are often found crystallized in bands. Large, nearly perfect specimens are occasionally found. These phenomena can be simulated by preparing a gel from water glass (a solution of Na_2SiO_3) to represent the silicate rock. Small quantities of other compounds are introduced, and crystals are formed by precipitation or oxidation reduction reactions. A solution of one of the ions of the desired crystal is added to an acidified solution of water glass. This mixture is allowed to stand. When the other reactant is added to the top of the solution the ions diffuse downward, and the reaction takes place. The longer these crystals are permitted to grow, the more beautiful they become.

Materials

Beaker, 400-mL
4 Test tube, 16- x 150-mm
4 Stoppers
Parafilm™
1.0 M Copper(II) chloride, $CuCl_2$, 100 mL (17.0 g $CuCl_2 \cdot 2H_2O$ per 100 mL solution)
1.0 M Lead acetate, $Pb(C_2H_3O_2)_2$, 100 mL (32.5 g $Pb(C_2H_3O_2)_2$ per 100 mL solution)
1.0 M Potassium iodide, KI, 100 mL (16.6 g KI per 100 mL solution)
1.0 M Copper(II) sulfate, $CuSO_4$, 100 mL (25.0 g $CuSO_4 \cdot 5H_2O$ per 100 mL solution)
1.0 M Sodium chromate, 100 mL (23.4 g $Na_2CrO_4 \cdot 4\,H_2O$ per 100 mL solution)
1 Iron nail
1 Zinc strip
1.0 M Acetic acid, $HC_2H_3O_2$, 100 mL (mix 6 mL glacial acetic acid per 100 mL solution)
Hot plate
Sodium silicate solution, 100 mL (15 mL of saturated solution diluted to 100 mL). The saturated solution can be purchased in hobby shops or pharmacies as water glass. If not available locally, Aldrich Chemical Co., (800) 558-9160, will send 1 liter (Cat. No. 33, 844-3) for $12.60 plus shipping charges.

Safety

Dispose of solutions according to local codes.

Procedure

Preparation of acidified gel.

1. Prepare a hot water bath by half filling a 400-mL beaker and placing it on a hot plate.
2. Put 25 mL 1.0 M acetic acid in a test tube. *(NOTE: The concentration of acetic acid and the sodium silicate solution are critical for gelling.)*
3. Add the specified amount of the first reactant (see following table) to the acetic acid in the test tube, and mix well. This may be done by stoppering the test tube and gently turning it upside down several times.
4. Add 25 mL sodium silicate solution to the acid mixture. Mix well as above, and cover with Parafilm (with a pinhole) or a stopper.
5. The gel will set in minutes if put into very hot water.
6. Add the metal or solution listed for specific systems (see following table), stopper, and display.
7. Observe daily.

Specific Systems:

To produce	First reactant (Step 3)	Then add (Step 6)
Copper tree	5 mL 1.0 M $CuCl_2$	Iron nail + 1.0 mL 1.0 M NaCl
Lead tree	1 mL 1.0 M $Pb(C_2H_3O_2)_2$	Zn strip + 4 mL water
Lead iodide	4 mL 1.0 M $Pb(C_2H_3O_2)_2$	10 mL 1.0 M KI
Copper chromate	4 mL 1.0 M Na_2CrO_4	4 mL 1.0 M $CuSO_4$

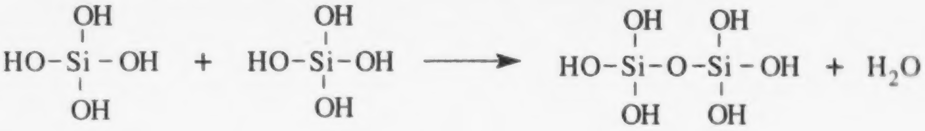
Demonstration 2: An Inorganic Polymerization Demonstration

Description

Silicates polymerize in acid solution according to the following equation:

$$Na_2SiO_3(aq) + 2HCl(aq) + H_2O(l) \rightarrow Si(OH)_4(aq) + 2NaCl(aq)$$

Although the product is represented by the formula $Si(OH)_4$, it consists of a complex mixture of polymeric acids formed by condensation.



$$HO{-}\overset{\displaystyle OH}{\underset{\displaystyle OH}{Si}}{-}OH + HO{-}\overset{\displaystyle OH}{\underset{\displaystyle OH}{Si}}{-}OH \longrightarrow HO{-}\overset{\displaystyle OH}{\underset{\displaystyle OH}{Si}}{-}O{-}\overset{\displaystyle OH}{\underset{\displaystyle OH}{Si}}{-}OH + H_2O$$

The $Si(OH)_4$ cannot be isolated from the mixture. If the gelatinous mixture is heated, the reaction is reversible. If the $NaCl(aq)$ is removed by washing, the reaction is not reversible. On heating, the unwashed silica becomes a solid, and the washed product is a powder that can absorb a large amount of water.

Materials

1 M Hydrochloric acid, HCl, 100 mL (8.4 mL conc. HCl diluted to 100 mL. Be sure to *add acid to water.*)
17% Sodium metasilicate, Na_2SiO_3, 100 mL (water glass, described in demonstration. 17 g Na_2SiO_3 is mixed with 83 g H_2O.)
Beaker, 250-mL
2 Petri dishes
Drying oven or microwave oven
0.05 M Silver nitrate, $AgNO_3$, 10 drops (0.02 g in 2 mL distilled H_2O)
Distilled or deionized water

Procedure

1. Add 100 mL 1 M HCl to 100 mL 17% Na_2SiO_3 in a weighed 250-mL beaker. Decant supernatant liquid.
2. Find the mass of the gel. Divide the gel into 2 equal parts.
3. Place half in a weighed Petri dish, and weigh again.
4. Wash the remaining gel in the beaker with distilled water until no cloudiness appears when a few drops of $AgNO_3$ is added to the decanted wash water.
5. Heat both gels in a drying oven at 105 °C for 1.5 hr or in a microwave oven for 15 min.
6. Reweigh the samples.

Data Analysis

1. Describe the differences in the two samples.
2. Students can calculate the number of grams of water absorbed by one gram of gel for each sample.
3. Explain the differences in structure that can account for the results in Question 2.
4. Which of the samples would be most likely to be packed with cameras and electronic equipment for use as a drying agent?

Demonstration 3: Supersaturated Solution

Description

The action of heat and pressure on water beneath the surface of the earth sometimes creates supersaturated solutions. This demonstration shows how such solutions could produce crystals.

Materials

Sodium acetate trihydrate, $NaC_2H_3O_2 \cdot 3H_2O$, 160 g
Distilled water, 30 mL
Tap water for water bath
Erlenmeyer flask, 500-mL
Beaker, 1-L or greater for water bath
Parafilm™ or 100-mL beaker to cover the Erlenmeyer flask
Graduated cylinder, 50- or 100-mL
Glass stirring rod
Wash bottle with distilled water
Balance, triple beam
Hot plate, laboratory burner or alcohol burner
Ringstand set up, if using burner
Heat-resistant gloves or tongs
Goggles and apron

Procedure

1. Weigh 160 g of sodium acetate trihydrate in a 500-mL Erlenmeyer flask.
2. Add 30 mL distilled water to the flask.
3. Heat the mixture in a hot water bath stirring occasionally until all of the solid is dissolved. (This may take 15 min or so.)
4. Remove all crystals from the sides of the flask by rinsing them down with small squirts of water from the wash bottle.
5. Cover the flask with Parafilm or the inverted 100-mL beaker.
6. Allow the solution to cool to room temperature undisturbed, or to speed up the cooling process, run cold water over the sides of the flask making sure no tap water gets into the flask and contaminates the solution.
7. While holding a single crystal of sodium acetate over the open mouth of the flask, snap your fingers, and allow the crystal to drop into the flask. The single crystal should start crystallization.
8. After crystallization is complete turn the flask upside down, and nothing should fall out.
9. The sides of the flask should be warm since this is an exothermic process.
10. The solution may be used over and over again by reheating it to redissolve the sodium acetate.

Remarks

The addition of too much water will result in leftover liquid after recrystallization.

Variations of this demonstration include placing a single crystal of sodium acetate in a shallow container and pouring the solution described above on the crystal to produce a "stalagmite" of sodium acetate. A buret may also be used to add the solution to the crystal.

Demonstration 4: Precipitation and Redissolution of Calcium Carbonate

Description

$CaCO_3$ is the major chemical component of many minerals and natural products. These minerals include marble, calcite, aragonite, coral, sea shells, stalactites, stalagmites. Precipitation of $CaCO_3$ forms these substances. Dissolution of $CaCO_3$ allows the raw materials to be transported.

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightleftharpoons CaCO_3(s)$$

This demonstration illustrates how sensitive the direction of this equilibrium expression is to reaction conditions (pH, carbonate concentration, temperature).

Materials

1 L Saturated $Ca(OH)_2$ (limewater) [Boil 1 L distilled water, cover, and allow to cool overnight. This removes CO_2 from the water. Add 1.8 g $Ca(OH)_2$, mix well, and allow to settle. Filter if cloudy at the time of use.]
Carbon dioxide, CO_2 (either dry ice (solid CO_2) or a cylinder of compressed CO_2)
4 Straws (or 25-cm lengths of 6- to 8-mm glass tubing, fire polished)
Gloves or towels to handle dry ice
Hot plate with magnetic stirrer and stirring bar
2 Beakers, 500-mL

Safety

The $Ca(OH)_2$ solution is basic with a pH = 12.4. It can irritate eyes and skin. In case of ingestion, give copious drinks of water to dilute the limewater in the stomach, and seek medical advice.

Procedure

1. Place 150-250 mL of clear, saturated $Ca(OH)_2$ in a 500-mL beaker. Have a volunteer blow gently through a straw or glass tube into the solution. Caution: because $Ca(OH)_2$ is a strong base, warn the volunteer to not get any of the liquid into the mouth (see *Safety* above). A white precipitate of $CaCO_3$ will form.
2. Place a second sample of $Ca(OH)_2$ in a second beaker. Bubble pure CO_2 through the solution by dropping in a piece of dry ice or by using the compressed gas. Initially, a white precipitate of $CaCO_3$ will form. However, the precipitate will slowly dissolve eventually giving a clear solution.
3. Have your volunteer return to try to redissolve the first sample of $CaCO_3$ by blowing again into the sample. No amount of blowing will redissolve the precipitate.
4. Put the solution from Step 2 on a hot plate with stirring. The carbon dioxide gas is less soluble in hot water and is removed. Near the boiling temperature, the solid begins to reform. When the solution cools, the appearance of suspended solid will be quite pronounced. (Adding pure CO_2 to this final solution will not redissolve the $CaCO_3$.)

Explanations

Step 1. Although H_2CO_3 is not stable, it is useful to view CO_2 dissolving in water to form carbonic acid.

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq) \qquad (a)$$

The acid is immediately neutralized by hydroxide forming carbonate ion that in turn reacts with the calcium ion forming insoluble $CaCO_3$.

$$H_2CO_3(aq) + 2OH^-(aq) \rightleftharpoons 2H_2O(l) + CO_3^{2-}(aq) \qquad (b)$$

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightleftharpoons CaCO_3(s) \qquad (c)$$

Step 2. When pure CO_2 is being added, its concentration increases until all the hydroxide originally in the solution is consumed. Thus, any excess dissolved CO_2 is unneutralized carbonic acid. Some hydrogen ions dissociate and dissolve the precipitate converting the carbonate to bicarbonate. ($CaHCO_3$ is soluble.)

$$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq) \qquad (d)$$

$$H^+(aq) + CaCO_3(s) \rightleftharpoons Ca^{2+} + HCO_3^-(aq) \qquad (e)$$

Step 3. Exhaled breath is only 4% CO_2. By contrast, the gas obtained from dry ice is 100% CO_2. Blowing into the solution never provides a concentration of carbonic acid high enough to exceed the available hydroxide. Think of all the reactions as reversible. Thus, in each, the reactants and products are in a state of balance. When the pressure of $CO_2(g)$ in reaction (a) cannot get very high, reaction (b) cannot be forced to consume all the OH^-. Thus the effect of the last two reactions cannot be seen. That is, the equilibrium in reaction (d) and (e) lies far to the left.

Step 4. Heating the solution decreases the solubility of CO_2 in water. Thus reaction (a) is forced to the left. This in turn forces reactions (d) then (e) to the left and $CaCO_3$ reforms.

Demonstration 5: Pot-O-Gold

Purpose

This demonstration should precede a discussion on solubility. Solubility and precipitation are important in understanding the occurrence of minerals. This demonstration catches the attention of students who watch the platelets swirling and catching the light.

Materials

Nonconsumables

Beaker, 1-L
Florence flask, 500-mL, with stopper
Burner
Funnel

Consumables

0.20 M Lead nitrate, $Pb(NO_3)_2$, 50 mL (3.3 g $Pb(NO_3)_2$ per 50 mL solution)
0.20 M Sodium iodide, NaI, 50 mL (1.5 g NaI per 50 mL solution)
Filter paper

Safety

Both lead and iodide ions are listed as hazards. Students should not handle these chemicals. Normal precautions should be followed. Excess solid lead iodide should be stored in a lead-waste container. The filtrate is used as part of the second solution.

Procedure

1. Mix the lead nitrate and sodium iodide solutions. A yellow precipitate of lead iodide will form.
2. Filter the supernatant liquid into 1-L beaker and add enough distilled water for a total volume of 300 mL.

3. Heat solution containing the filtrate to boiling.
4. Add enough solid lead iodide to make a saturated solution at 100 °C.
5. Pour the saturated solution, while hot, into a 500-mL Florence flask.
6. Let the solution cool slowly to room temperature. The lead iodide will precipitate out as shiny platelets that look like gold.
7. Stopper the flask securely. The mixture can be kept in a stoppered flask for several years.

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. What happens on a particle level when crystals form from evaporation of a saturated solution? *[As water slowly evaporates, the ions deposit on the crystal because the saturated solution cannot dissolve more solute.]*
2. How can flame tests and borax bead tests be used in mineral identification? What are the limitations of these tests? *[Both flame tests and borax bead tests are used to identify the positive ion of the mineral. Sometimes the colors are very similar and these tests cannot identify the negative ion.]*
3. What is the basic unit common to all silicate minerals? What is its geometry? *[The SiO_4^{4-} ion. It is tetrahedral.]*
4. How are crystals formed from a supersaturated solution? *[A single crystal is added to disturb the solution and then many crystals form rapidly.]*
5. What factors are characteristic of minerals that are classified as gemstones? *[Color, transparency, and luster or brilliance.]*
6. Compounds of the transition elements often exhibit color. To what can this color be attributed? *[If the transition element forms a compound in which its d orbitals do not all have the same energy, the result is an energy gap that matches the energy of photons of visible light. When white light strikes the compounds, those photons are absorbed. The color we see is the result of the colors that are not absorbed.]* (See *Transition Elements* module.)
7. How is a gemstone different from a mineral? *[A gemstone, generally, forms when a particular transition element atom is incorporated as an impurity in a parent mineral. For example, ruby results from Cr^{3+} ions replacing some of the Al^{3+} ions in the mineral corundum, Al_2O_3.]*

Counterintuitive Examples

Salol (phenyl salicylate), a solid at room temperature, is a substance that supercools readily. Supercooling is the process of cooling a liquid below its freezing point without changing to a solid. Prepare microscope slides that have two "globs" of salol as shown here. Have students melt the larger glob by holding a match under it. After melting the glob place the slide on the desk top to allow it to cool to room temperature. Ask how the solid and liquid can both exist at the same temperature. Then use a spatula or stirring rod to break off a crystal from the smaller glob and move it to the puddle formed from the larger glob. The slow crystallization holds the interest of students.

Analogies and Metaphors

Fruit is frequently stacked in supermarkets in a pattern analogous to cubic close packing.

Pictures in the Mind

Making Models of Silicate Minerals

Introduction

Students sometimes have great difficulty visualizing in three dimensions. This activity gives them hands-on experience imagining various silicate ions. Commercial chemistry models can be used by declaring that the ball that usually represents the carbon atom represents a silicon atom. However, gum drops and tooth picks are better because students will have to decide the appropriate angles.

Materials

Model kits or gum drops of two colors and tooth picks.

Procedure

1. Require students to work in groups of two or three.
2. Explain that silicon has four covalent bonds and the bonds need to be as far apart as possible. Remind them to consider three dimensions.
3. Direct students to make an SiO_4^{4-} ion, and draw a projection on paper. Circulate through the room correcting those insisting on using only two dimensions.
4. Make models and drawings of:
   - a. $Si_2O_7^{6-}$
   - b. chains of $(SiO_3^{2-})_n$ and double chains of $(Si_4O_{11}^{6-})_n$
   - c. planar sheet of atoms where each Si atom shares oxygen atoms with three other silicons
   - d. a three-dimensional structure of empirical formula SiO_2.
5. Have students describe the properties of minerals that would have each of the structures they made. *Tips for the Teacher* and the *Transparency Masters* will be useful in interpreting students' drawings and providing examples. If samples of minerals with each type of structure—chains, sheets, and three-dimensional—are available, they will add to the understanding of the relationship between structure and properties.

TIPS FOR THE TEACHER

Language of Chemistry

crystal substance in which the atoms, ions, or molecules are arranged in an orderly, repeating three-dimensional pattern called a crystal lattice.

lithosphere the upper portion of the earth's crust.

mineral inorganic solid of definite composition found in the earth's crust.

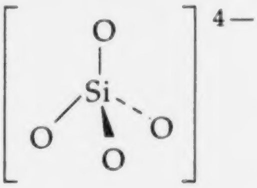
silicates minerals with crystal structure containing silicon-oxygen tetrahedra.

supercooling process of cooling a liquid below its freezing point without changing to a solid.

Silica and **silicates** are compounds made primarily of silicon and oxygen atoms. They can be discussed to make three important points:

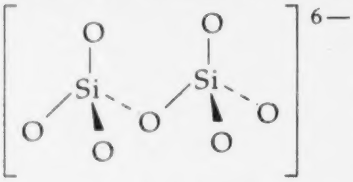
1. Many minerals are silicate based.
2. Lewis structure pictures can be used to classify minerals.
3. Lewis structure pictures can help explain physical properties of minerals.

All silicate minerals contain Si atoms bonded to four O atoms in a tetrahedral arrangement. The simplest involve the orthosilicate ion (SiO_4^{4-}).



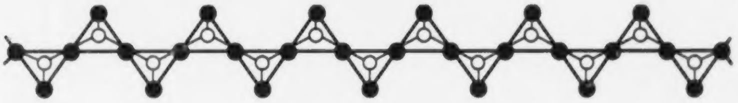
Each oxygen is bonded to only one Si and has a residual negative charge. Examples of orthosilicate minerals are zircon ($ZrSiO_4$) and forsterite (Mg_2SiO_4).

When an oxygen atom is shared by two Si atoms, the pyrosilicate anion ($Si_2O_7^{6-}$) is obtained.

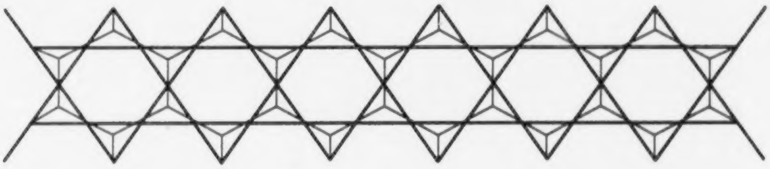


The Lewis structure shows single bonds, octets on each atom, and a tetrahedral geometry around the Si atom. Some pyrosilicate minerals are akermanite ($Ca_2MgSi_2O_7$), lawsonite ($CaAl_2Si_2O_7(OH)_2$), and hemimorphite ($Zn_4(OH)_2Si_2O_7$).

The silicate units can continue to link through sharing oxygen atoms to form long single chains of the general formula SiO_3^{2-}

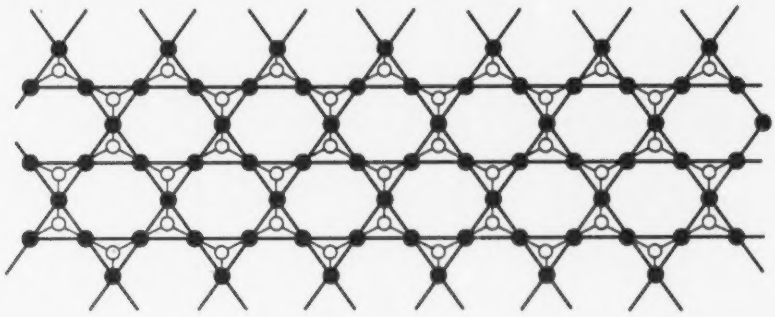


or long double chains having the formula $Si_4O_{11}^{6-}$.



Examples of some of SiO_3^{2-} minerals are enstatite ($MgSiO_3$) and spodumene ($LiAl(SiO_3)_2$). Tremolite ($Ca_2Mg_5Si_8O_{22}(OH)_2$) is an example of a mineral with double chains.

SiO_x tetrahedra can also link into two-dimensional planar sheets of atoms, where each Si atom shares oxygen atoms with three other silicons. These minerals have the general formula $Si_4O_{10}^{4-}$ and are called sheet silicates. Chrysolite ($Mg_6Si_4O_{10}(OH)_8$) is an example. In micas one Si is replaced by an Al, for example, muscovite ($KAl_2AlSi_3O_{10}(OH)_2$).



A logical extension of this build-up of tetrahedra is to have all four oxygens in a tetrahedral silicate unit shared by Si atoms in adjacent units. The result is quartz, with an empirical formula of SiO_2. The quartz structure can be imagined by viewing the three-dimensional linked structure of diamond. SiO_2 would have Si's where each C is in the diamond structure; the oxygen atoms are then midway between each pair of Si atoms.

The one-, two-, or three-dimensional links of SiO_x units lead to their physical properties. Asbestos is mixed minerals of the long, double-chain type. Thus, they are needle-like, which may account for some of the damage they do to living tissue. Micas are slippery due to two-dimensional sheets of SiO_x units that can slide over each other. Quartz has structural strength because of its three-dimensional covalent linking of SiO_x units.

Problem Solving

Many common minerals contain oxygen as a major component. Some of these are easily collected and identified. This problem-solving activity shows the variation of that oxygen content. Problem: Calculate the oxygen content of flint or quartz (SiO_2).

Find the molar mass of the compound:

$$\begin{aligned} 1 \text{ mol Si} &= 28.1 \text{ g} \\ 2 \text{ mol O} &= \underline{32.0} \text{ g} \\ \text{molar mass} &= 60.1 \text{ g} \end{aligned}$$

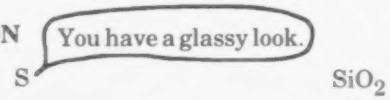
The part of the molar mass due to oxygen can be calculated as a percentage of the whole:

Sample Problems: Calculate the percent oxygen in the minerals:

1. Calcite, $CaCO_3$ *[48.0%]*
2. Corundum, Al_2O_3 *[47.1%]*
3. Gypsum, $CaSO_4 \cdot 2H_2O$ *[59.2%]*
4. Hematite, Fe_2O_3 *[30.1%]*
5. Orthoclase (K-feldspar), $K(AlSi_3O_8)$ *[50.1%]*

HUMOR: ON THE FUN SIDE

1. **CHEMTOON**


You have a glassy look.
S
SiO_2

Chem 13 NEWS, February 1985, p. 16

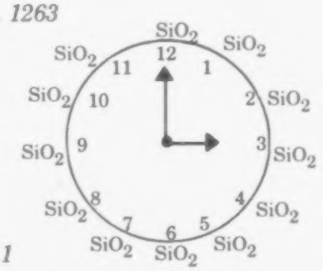
2. **MNEMONIC DEVICE**

Osaic—The five most abundant crustal elements of the earth (oxygen, sulfur, aluminum, iron, calcium.

CHEM 13 NEWS, November 1977, p. 1263

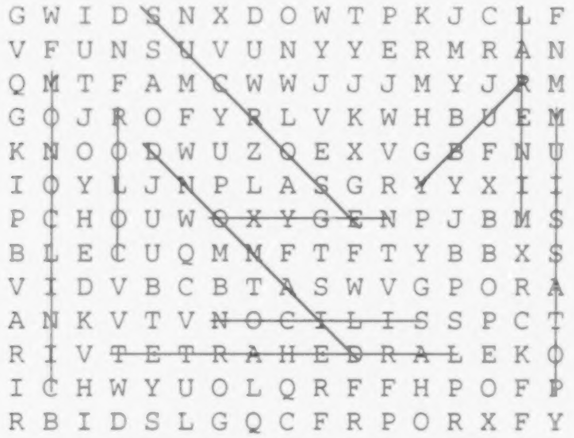
3. This is a HARD one! What is it?

Answer: Rock around the clock!



CHEM 13 NEWS, February 1982, p. 1

4. Word Search (see *Appendix* for master copy)



G W I D S N X D O W T P K J C L F
V F U N S U V U N Y Y E R M R A N
Q M T F A M C W W J J J M Y J R M
G O J R O F Y R L V K W H B U E M
K N O O D W U Z O E X V G B F N U
I O Y L J N P L A S G R Y Y X I I
P C H O U W ~~O X Y G E N~~ P J B M S
B L E C U Q M M F T F T Y B B X S
V I D V B C B T A S W V G P O R A
A N K V T V ~~N O C I L I S~~ S P C T
R I V ~~T E T R A H E D R A L~~ E K O
I C H W Y U O L Q R F F H P O F P
R B I D S L G Q C F R P O R X F Y

Words about concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Silicates are compounds made primarily of oxygen and this element.
2. A gem whose basic mineral is Al_2O_3 with Cr(III) as an impurity.
3. This property of gemstones is characterized by presence of transition elements as impurities.
4. This gemstone is an allotrope of carbon.
5. Most abundant element in the earth's crust, by mass.
6. Inorganic solid of definite composition found in the crust.
7. Silicates have this geometrical arrangement.
8. Rock candy is this organic compound.
9. One of the six basic crystal systems.
10. A lavender flame test identifies this element.

Answers: 1. SILICON 2. RUBY 3. COLOR 4. DIAMOND 5. OXYGEN 6. MINERAL 7. TETRAHEDRAL 8. SUCROSE 9. MONOCLINIC 10. POTASSIUM

5. See cartoons at end of module.

MEDIA

1. The *World of Chemistry* (high school version) videotapes. WINGS for Learning/ SUNBURST, 101 Castleton Street, Pleasantville, NY 10570; (800) 321-7511; (914) 747-3310; (914) 747-4109 (FAX).
   a. "The Chemistry of Earth": reinforcing chemistry topics through examples from geology—solubility, precipitation, equilibrium, and acid-base chemistry applied in explaining mineral deposition and world-wide distribution of ores and silicate minerals; the relationship between chemical structure and macroscopic properties—cave formation and stalactite and stalagmite development; limestone.
   b. "The Atmosphere": Common theory on formation of the atmosphere; the composition today—highlights of some environmental concerns: contaminants being added to the atmosphere, the greenhouse effect, the hole in the ozone layer, the commercial use of CFC's.
   c. "Chemical Bonds": Ionic and covalent bonds defined through graphic illustration—

a comparison of their relative strength in ionic crystals (sodium chloride) and network covalent solids (diamond)—molecular solids; the difference between the strong intramolecular covalent bonds between atoms in a molecule and the weaker intermolecular attractions, such as hydrogen bonds, between molecules.

2. MECC (Minnesota Educational Computer Consortium, St. Paul, MN 55165): "Murphy's Minerals," a computer game to practice skills learned in identifying minerals. (612) 569-1500.

3. Planet Earth Series: A seven-part video series that was first shown on the Public Broadcasting Service in 1986. This series is a rich and exciting investigation of the earth sciences today, a vivid panorama of the planet we call home. Produced by WQED/Pittsburgh, in association with the National Academy of Sciences, this series includes the episodes: "The Living Machine," "The Blue Planet," "The Climate Puzzle," "Tales from Other Worlds," "Gifts from the Earth," "The Solar Sea," and "Fate of the Earth."

   The episode most useful with this topic is "Gifts from the Earth." This program acquaints students with the breadth and variety of the world's natural resources through a story spanning millions of years. Included are segments explaining how mineral-saturated hydrothermal vents on the sea floor ultimately became gold and copper mines on land; about the slow accumulation of fossil fuels; of the ingenious scientists who seek out the secrets of the earth. The formation of metallic elements, minerals—both precious and common—are explored, as well as how these resources influence civilization. One of the handouts, "Minerals in Our Lives," is particularly interesting as it lists 60 common elements and compounds readily available from the earth's crust and describes uses of each. Available from: PBS Video, 1320 Braddock Place, Alexandria, VA 22314. Telephone: (703) 739-5000.

   The episode "The Blue Planet" deals primarily with the influence of liquid water on the surface of the earth, and could also be used in this module.

4. "Gemstones of America," STS Film & Video Productions, P.O. Box 27477, Salt Lake City, UT 84127. A Smithsonian Project, 1991. (801) 263-3959.

5. Software published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   a. *KC? Discoverer with Knowledgeable Counselor*, by Daniel Cabrol, John W. Moore and Robert C. Rittenhouse. Special Issue 2, for IBM PS/2, PC compatible computers.

   b. *KC? Discoverer: Exploring the Properties of the Chemical Elements*, by Aw Feng and John W. Moore. Vol. I B, No. 1, for IBM PS/2, PC compatible computers.

   c. *KC? Discoverer?*, by Michael Liebl, Vol. IV A, No. 2, for all Apple II computers.

   d. *The Periodic Table Stack*, by Michael Farris. Vol. I C, No. 1, for the Apple Macintosh.

6. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).

   a. For the Apple II computer: AP 807

   b. For IBM PS/2 PC-compatible computers: PC 3702

5. Videodiscs published by*JCE:* Software, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   a. "Earth's Atmosphere," "Acid Rain and Limestone" and "Silicates," three chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc II (double sided, 60 min.), Special Issue 4.

   b. *The Periodic Table Videodisc* (single side, 30 min.). Special Issue 1.

6. *Gems and Minerals: The Ultimate Rock Video*, available from Smithsonian Laserdisc Collection, Smithsonian Institution, Washington, DC; (202) 357-1300.

Links/Connections

WITHIN CHEMISTRY

Chemical bonding. The study of rocks and minerals reinforces the importance of structure and geometry in determining properties of network substances.

Equilibrium. Equilibrium can be reviewed through study of the crystallization process.

Solubility and precipitation. Growing crystals, particularly in gels, is a vivid illustration of properties of solutions and precipitation phenomena.

Transition elements. Transition elements are frequently responsible for color that makes minerals valuable as gemstones.

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

There is a strong connection between the topic of this module and geology.

Gemstones

Gemstones are minerals valued for their beauty including color, transparency, and luster or brilliance. Color is sometimes dependent on crystal structure and bond type, but it most frequently is related to chemical composition. Color is often characterized by the presence of transition elements either directly related to the composition, such as the blue or green of copper-containing minerals or as impurities that impart color to colorless minerals such as quartz.

The basic structures of minerals are frequently silicon oxides, which have already been discussed, or a form of Al_2O_3, corundum. It is a hexagonal closest packed arrangement of O^{2-} ions with Al^{3+} ions in two-thirds of the octahedral holes. The oxide ions are stacked in the A-B-A-B arrangement, and the aluminum ions fall into those holes that allow them to touch six oxide anions. Examples of the gemstones formed are shown below.

Gem	Color	Basic Mineral	Impurity
Ruby	Red	Al_2O_3	Cr^{3+}
Amethyst	Purple	SiO_2	Mn^{3+} or Ti^{2+}
Topaz	Yellow	$Al_2SiO_4(F,OH)_2$	
Blue sapphire	Blue	Al_2O_3	V^{3+} or Co^{3+}
Emerald	Green	$Be_3Al_2Si_6O_{18}$	
Opal	White or colorless	$SiO_2 \cdot nH_2O$	
Jade	Green	$Ca_2Si_8O_{22}(OH)_2$	Mg^{2+} or Fe^{2+}

TO THE CONTEMPORARY WORLD

Personal (including careers)

Study of rocks, minerals and gems will be interesting to those who are interested in careers as geochemists as well as those who pursue "rock hounding" as a hobby.

One area that will almost certainly be emphasized in the future is materials science—extracting and modifying earth materials for practical uses (see *Materials Science* module).

Community

Field Trips

In areas where possible, a trip to a working mine would be valuable. Many cities have rock and mineral displays in museums.

Knowledgeable Individuals

The earth science or geology teacher or local college professor
Rock hobbyists
Members of the local geological society
Lapidarist—a collector or dealer in gems

Other

1. A little item that can be purchased in a toy store and uses the crystallization concept is called the Magic Tree. It uses a cardboard three-dimensional tree and a powdery compound (like copper sulfate) to make crystals grow from an evaporating solution. It is manufactured by New Tomorrow, 7251 Garden Grove Blvd. #E, Garden Grove, CA 92641.
2. An article in *ChemMatters* (published by the Education Division of the American Chemical Society) entitled "Growing Diamonds" (April, 1990, pp. 10-13) describes the crystalline nature of diamond as it relates to graphite and presents some interesting comparisons between the two. In this same issue, the feature "Back Burner" also uses diamonds as its theme.
3. An essay in Aldo Leopold's Sand County Almanac (1966) New York: Oxford University Press is especially fascinating as it describes the way in which an atom, called "X," cycles through the biosphere as a part of a much larger mineral cycle. The essay "Odyssey" is found in Part 2 (pp. 104-108).
4. The October 5, 1990 issue of *Science* (pp. 25-26) has a interesting article about the production of isotopically pure diamond—the best room temperature heat conductor ever made. Controversy between an amateur scientist and General Electric about the origin of the idea is described.
5. Mica is a sheet silicate that can be cleaved to make large translucent sheets. These sheets were used as window closures prior to the advent of plate glass.
6. The brilliance of gemstones depends largely on their high refractive index. Because diamond's refractive index is so large, light rays exhibit such small angles of refraction that the rays often fail to leave the material, but will be reflected back into it. This phenomenon is known as total reflection. It can be achieved in a diamond if the stone is properly cut. When a diamond exhibits total reflection, none of the light entering the face of the stone will be able to escape out the back, but will be totally reflected back to the face, imparting a brilliance and sparkle to diamond that has only been matched recently by synthetic gems of almost equally high refractive index.

Extensions

1. Long-term growth of crystals and comparison of crystal shapes.
2. Have students do a written or oral report on their birthstone. Identify its composition and the origin of the stone's color (see *References*). Discuss the crystal system to which the stone belongs (see Figure 4). (Idea from Angie Matamoros.)
3. Concrete can be made in the laboratory and provides excellent practice in controlling variables. It is, however, a messy procedure, and students must be cautioned about spilling the ingredients on equipment and clothing and about pouring cement containing wastes in sinks. A convenient ratio is 14.5 g cement, 41 g sand, 57 g gravel, and 11.5 g water. Use disposable containers such as paper cups or coffee cans. The materials must be stirred while mixing and allowed to stand several days. The concrete can be analyzed, and other combinations may be tried. More complete instructions are available in the written materials accompanying "Chemical Kinetics" (Program 7) The World of Chemistry, High School Version, Wings for Learning, 1600 Green Hills Road, P.O. Box 660002, Scotts Valley, CA 95067-0002.
4. **A Chemist's Map of the United States**

   Examine "A Chemist's Map of the United States" (see the transparency master in the *Appendix*). List all cities that describe the solid state.
5. **A Chemist's Map of Your State**

   Refer to a detailed Atlas of the United States, selecting your state map. Find and make a list of site names associated with chemistry. Prepare an outline sketch of your state map, situating the sites at appropriate points on the map.

References

Module drafted by Diana Doepken Rischling, Michael Pavelich, and Patricia Smith, the Colorado (South) team.

Aereszko, L. S. (1987). Rock candy in a cellophane bag. *Journal of Chemical Education*, *64*, 804.

Baer, C. D. (1990). The growth of large single crystals. *Journal of Chemical Education*, *67*(5), 410-412.

General instructions and discussion for easy growing of crystals. Tables of suggested chemicals and of crystal systems.

Bartlett, D., and Aguila, E. (1987). Teaching the concepts of metallurgy through the use of postage stamps. *Journal of Chemical Education*, *64*, 526-528.

Booth, M., and Lamb, W. (1983, October). Crystal growing [Experimenter's notebook]. *ChemMatters*, *1*(3), 12.

Detailed recipe for making Rock Candy.

Davenport, D. (1990). Burning diamonds and squeezing peanuts [Back burner]. *ChemMatters*, *(8)*2, 14-15.

A historical perspective on diamond synthesis.

Davenport, D. (1983, October). Brooklyn bridge, the structure of metals, and footprints in the sand [Back burner]. *ChemMatters*, *1*(3), 14-15.

A fascinating discussion of atomic packing in crystals and its consequences. This issue is filled with crystal articles.

Finkbeiner, A. (1984). Starborn: The origin of elements. *ChemMatters* *2*(3), 6-7.

A nice review of cosmology, tracing the nuclear reactions to convert hydrogen, eventually, into all of the elements.

Holden, A., and Singer, P. (1960). *Crystals and crystal growing*. Garden City, NY: Doubleday.

Holtzman, D. (1987, October). Superconductivity. *ChemMatters*, *5*(3), 18-21.

An introductory historical and theoretical discussion of superconductivity.

Hostettler, J. D. (1985). Geochemistry for chemists. *Journal of Chemical Education*, *62*, 823-831.

Kilday, C. (1987, October). Super-student conductors. *ChemMatters*, *5*(3), 22-23.

Report of and recipe for high school laboratory synthesis of $YBa_2Cu_3O_7$ superconductor.

Lamb, W. (1984, April). Liquid crystal displays. *ChemMatters*, *2*(2), 10-11.

A discussion of the design of LCD's and the molecule that makes them "tick."

Leopold, A. (1966). *Sand County almanac*. New York, NY: Ballantine Books.

Loeffler, B. M., and Burns, R. G. (1976, November-December). Shedding light on the color of gems and minerals. *American Scientist*, *64*, 636-647.

A discussion to how the absorption of light gives minerals their distinctive hues.

Marshall, E. (1990, October). GE's cool diamonds prompt warm words. *Science*, (250), 25-26.

Meadows, R. (1988, April). Fossil molecules. *ChemMatters, 6*(2), 4-7.

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Tanis, D. (1984, February). Underground sculpture. *ChemMatters, 2*(1), 10-11.

The chemistry and mechanism of stalactite formation is outlined.

Viola, V. E. (1990). Formation of the chemical elements and the evolution of our universe. *Journal of Chemical Education, 67,* 723–730.

Walker, L., and Lee, P. J. (1990). The rocky road to chemistry. *Journal of Chemical Education, 67,* 325–326.

Yamana, S. (1987). Cubic close packed structure model. *Journal of Chemical Education, 64,* 1040.

Zaug, H. (1990). Growing diamonds. *ChemMatters, (8)*2, 10-13.

Discusses the structure of diamond, the history of its synthesis and details of new techniques for low pressure growth.

Some periodicals of interest for rockhounds and earth science teachers (suggested by Leonard Himes):

Lapidary Journal. Devon Office Center, Suite 201, 60 Chestnut Avenue, Devon, PA 19333-1312.

General interest and hobbyist articles on jewelry, gems, minerals, fossils, and collecting localities. Calendar of mineral and gem shows (monthly).

Gems and gemology. 1660 Stewart Street, Santa Monica, CA 90404.

Discussion of state-of-the-art gemological techniques, new gemstones, and gem locality information. Written for the practicing gemologist. Many good semi-technical articles (quarterly).

Earth. 21027 Crossroads Circle, Waukesha, WI 53187.

General interest articles, written for the public, on structural and historical geology, oceanography, and meteorology as well as fossils, rocks, and minerals (monthly).

Colored stone. Devon Office Center, Suite 201, 60 Chestnut Avenue, Devon, PA 19333-1312.

Discusses new gemstones and identification problems, business trends, and reports. Some background articles on gemstone minerals and mining. Written for commercial jewelers (bimonthly).

The mineralogical record. P.O. Box 35565, Tucson, AZ 85740.

Mineralogy of worldwide localities, news of mineral collectors and museums, show reports, columns, and commentary. Written for advanced level collectors and mineralogists (bimonthly).

Rocks & minerals. 1319 Eighteenth Street, N.W., Washington, DC 20036-1802.

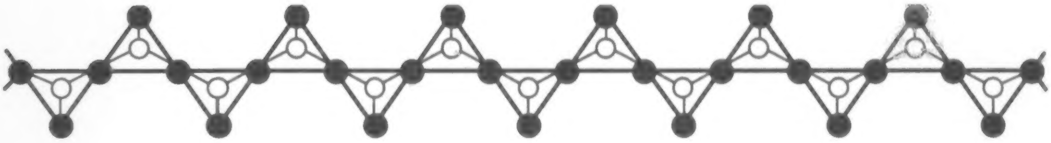
Mineralogy, paleontology, and geology oriented article and news information of interest to collectors. Includes columns and calendar of events. Education is a central theme of this publication (bimonthly).

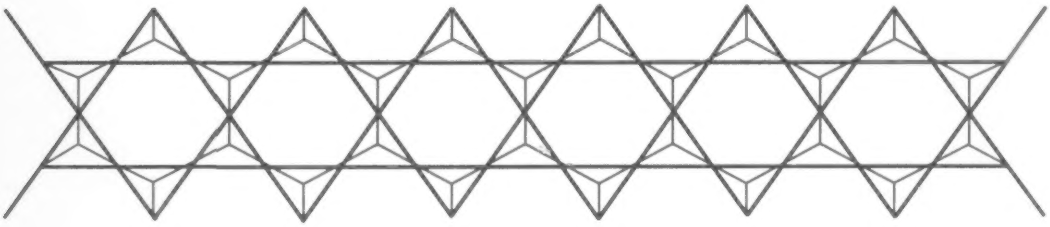
Needless to say, there are plenty of introductory texts on mineralogy and geology, all of which would be helpful to teachers. The magazines listed above also list and review books that are (often) less technical. There are many field guides (Audubon, Peterson, *etc.*) that are available in any mall bookstore for teachers who wish to enrich their basic knowledge.

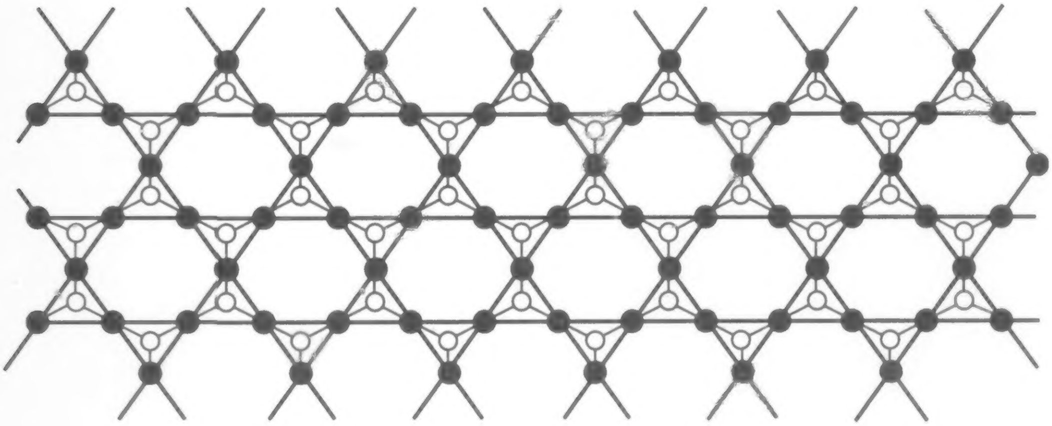
Appendix

- **Transparency Masters**
  1. Structures for SiO_3^{2-}, $Si_4O_{11}^{6-}$, and SiO_x
  2. A Geological Periodic Chart
  3. Word Search
- **Humor**

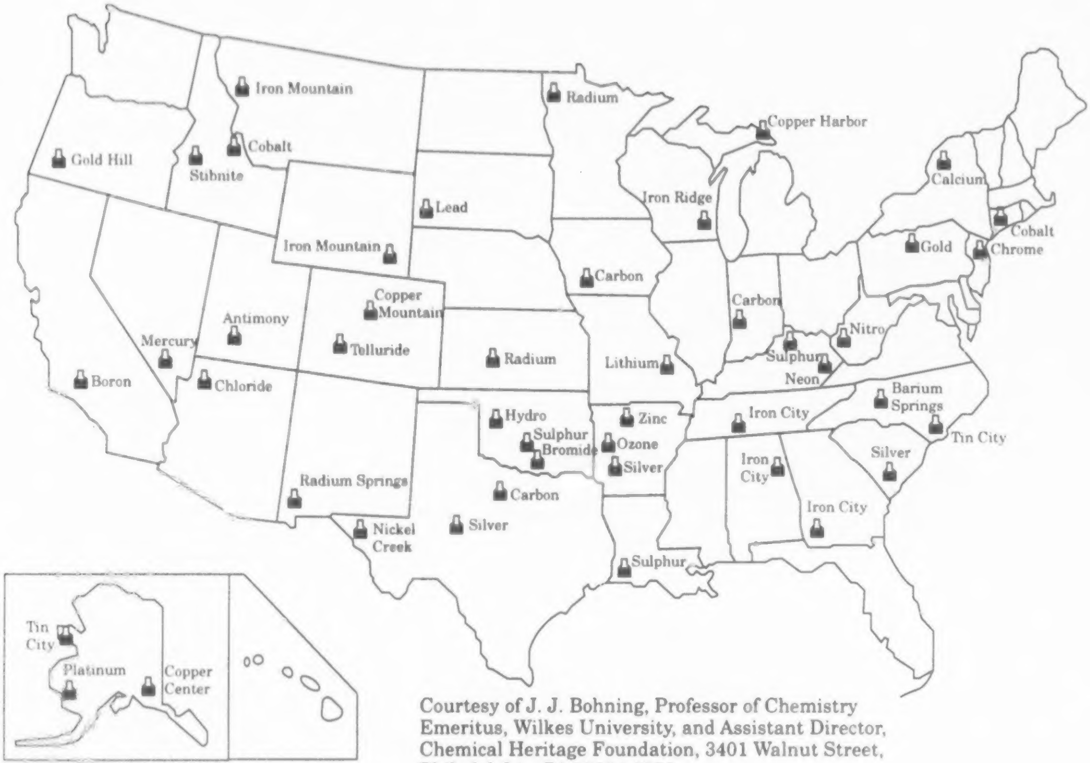
Structures for SiO_3^{2-}, $Si_4O_{11}^{6-}$, and SiO_x







A Geological Periodic Chart


Iron Mountain
Radium
Copper Harbor
Gold Hill
Stibnite
Cobalt
Calcium
Lead
Iron Ridge
Iron Mountain
Gold
Cobalt
Chrome
Carbon
Copper Mountain
Antimony
Carbon
Nitro
Mercury
Telluride
Radium
Lithium
Sulphur
Neon
Boron
Chloride
Barium Springs
Hydro
Zinc
Iron City
Tin City
Sulphur
Bromide
Ozone
Silver
Silver
Iron City
Radium Springs
Carbon
Iron City
Nickel Creek
Silver
Sulphur
Tin City
Platinum
Copper Center

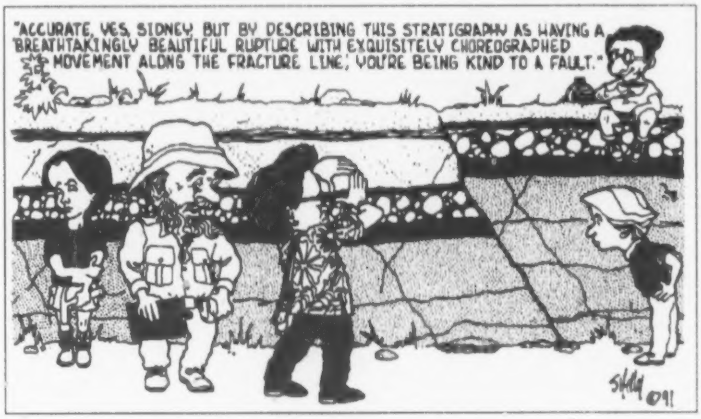
Courtesy of J. J. Bohning, Professor of Chemistry Emeritus, Wilkes University, and Assistant Director, Chemical Heritage Foundation, 3401 Walnut Street, Philadelphia, PA 19104-6228

Word Search

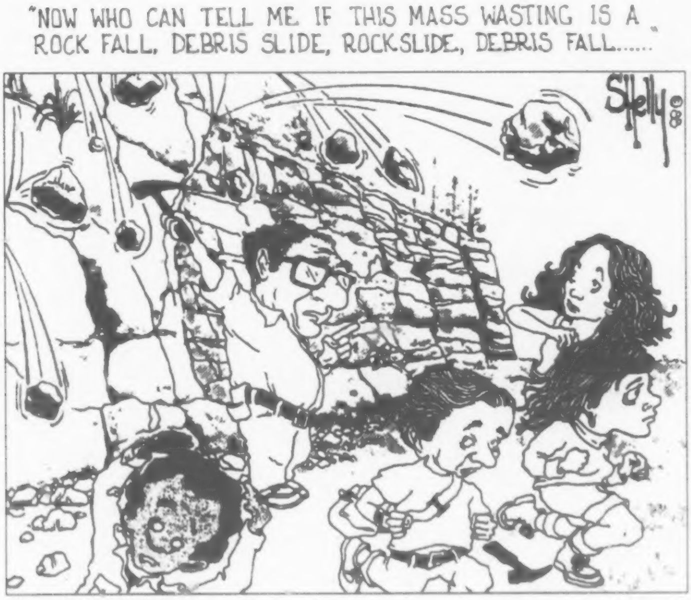
G	W	I	D	S	N	X	D	O	W	T	P	K	J	C	L	F
V	F	U	N	S	U	V	U	N	Y	Y	E	R	M	R	A	N
Q	M	T	F	A	M	C	W	W	J	J	J	M	Y	J	R	M
G	O	J	R	O	F	Y	R	L	V	K	W	H	B	U	E	M
K	N	O	O	D	W	U	Z	O	E	X	V	G	B	F	N	U
I	O	Y	L	J	N	P	L	A	S	G	R	Y	Y	X	I	I
P	C	H	O	U	W	O	X	Y	G	E	N	P	J	B	M	S
B	L	E	C	U	Q	M	M	F	T	F	T	Y	B	B	X	S
V	I	D	V	B	C	B	T	A	S	W	V	G	P	O	R	A
A	N	K	V	T	V	N	O	C	I	L	I	S	S	P	C	T
R	I	V	T	E	T	R	A	H	E	D	R	A	L	E	K	O
I	C	H	W	Y	U	O	L	Q	R	F	F	H	P	O	F	P
R	B	I	D	S	L	G	Q	C	F	R	P	O	R	X	F	Y

Words about concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Silicates are compounds made primarily of oxygen and this element.
2. A gem whose basic mineral is Al_2O_3 with Cr(III) as an impurity.
3. This property of gemstones is characterized by presence of transition elements as impurities.
4. This gemstone is an allotrope of carbon.
5. Most abundant element in the earth's crust, by mass.
6. Inorganic solid of definite composition found in the crust.
7. Silicates have this geometrical arrangement.
8. Rock candy is this organic compound.
9. One of the six basic crystal systems.
10. A lavender flame test identifies this element.


"ACCURATE, YES, SIDNEY, BUT BY DESCRIBING THIS STRATIGRAPHY AS HAVING A BREATHTAKINGLY BEAUTIFUL RUPTURE WITH EXQUISITELY CHOREOGRAPHED MOVEMENT ALONG THE FRACTURE LINE, YOU'RE BEING KIND TO A FAULT."

Cartoon by Shelly Fischman. Reprinted with permission.


"NOW WHO CAN TELL ME IF THIS MASS WASTING IS A ROCK FALL, DEBRIS SLIDE, ROCKSLIDE, DEBRIS FALL......"

Cartoon by Shelly Fischman. Reprinted with permission.

IN TODAY'S LAB CLASS, WE WILL EXAMINE AND MEMORIZE THE CHARACTERISTICS OF THESE 65 ROCKS!

Danylo Shepelavy, U Penn, 1993.
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A SourceBook Module

SIMPLE CHEMICAL REACTIONS

Version 2.0 1994

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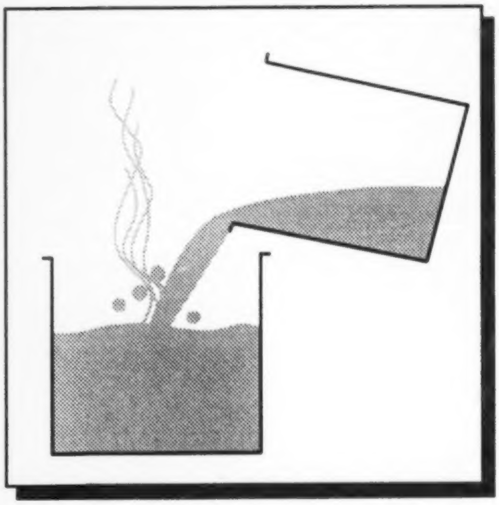

ChemSource

Instructional Resources for Preservice and
Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

The process by which one or more substances change into new substances is called a chemical reaction. This definition is a bit ambiguous and stuffy; chemical reactions are neither. When a reaction occurs it is usually obvious and often spectacular. Chemical reactions are the motion in our lives, and they are the heart of this science. Most of chemistry is devoted to learning more about these reactions—how much product they produce, how fast they proceed, and how much energy is involved. For students, chemical reactions supply the gee, the whiz-bang, the magic of chemistry. It is valuable for students to see many reactions *via* laboratory activities and teacher demonstrations, especially early in the course.

Chemical reactions involve the breaking and making of bonds, or stating it differently, a change in the location of electrons with respect to the nuclei of atoms. A driving force for reactions is that electrons find a lower energy environment, illustrated by the reaction of active metals in acid.

$$Mg(s) + 2\,HCl(aq) \rightarrow H_2(g) + MgCl_2(aq)$$

A major change here is that two of the electrons originally on the Mg atom have moved because the H nuclei can hold them more strongly. Chemical reactions are primarily movement of electrons so chemistry textbooks spend more time talking about electrons than they do about nuclei.

Breaking chemical bonds requires energy and making them produces energy so chemical reactions involve energy changes. Many times we do reactions to obtain the energy rather than the products of the reaction. The burning of fuels is an example. Several demonstrations in this module focus on the energy changes in chemical reactions.

Chemists use equations with symbols and formulas to represent the reactions that occur. It is critical that students understand that the equations we write on paper represent what happens in the laboratory. Balancing a chemical equation, in the early stages of the student's study, is best done by trial and error. Again, students should recognize that the numbers represent the quantities of substances that react in the laboratory.

The study of chemical reactions, for chemists and students, begins in the laboratory. Chemists mix substances together and look for evidence (such as formation of gas or precipitate, color change, or energy change) to indicate that a reaction has occurred. Additional experiments are conducted to identify the products. It is only through extensive experimentation that enough information is accumulated to enable us to predict the products of some reactions. Students should be encouraged not only to make the observations, but also to speculate about their meaning.

At first students may be bewildered by the wide variety of reactions that happen in the laboratory and the real world. They can gain some sense of control by developing and applying schemes for classifying reactions. Early in their exposure to chemistry they may find simple schemes such as synthesis, decomposition, replacement, and double replacement useful. As they learn more of the content of chemistry, their classification schemes will become more sophisticated and will include such concepts as acid-base and oxidation-reduction reactions.

PLACE IN THE CURRICULUM

The investigation of reactions as outlined in this module can come very early in the course. Laboratory activities and demonstrations create interest and provide an experiential basis for the study of later topics such as moles, atomic theory, electronic structure, and acids and bases. The topic of classification of reactions can be revisited later in the year when students can apply a greater knowledge base.

CENTRAL CONCEPTS

1. Chemical reactions are recognized by changes in the properties of the species involved.
2. Many chemical reactions can be classified, at the simplest level, according to general kinds of change involved:

$A + B \rightarrow AB$	Synthesis
$AB \rightarrow A + B$	Decomposition
$AB + C \rightarrow AC + B$	Replacement
$AB + CD \rightarrow AD + CB$	Double replacement

3. A chemical equation represents, in symbolic terms, the overall change in a chemical reaction at the atomic and molecular level (words, symbols, pictures in the mind).
4. All chemical equations must be "balanced" to be consistent with the Law of Conservation of Matter.

RELATED CONCEPTS

1. Chemical Formulas and Symbols
2. Law of Conservation of Mass
3. Elements, Compounds, Substances, Mixtures
4. States of Matter

RELATED SKILLS

1. Observe changes that occur in the laboratory
2. Use a balance
3. Light and use a burner safely
4. Recognize equation patterns

PERFORMANCE OBJECTIVES

After completing their study of chemical reactions, students should be able to:

1. write correct formulas given ion charges.
2. interpret formulas.
3. classify a given chemical equation as synthesis, decomposition, single replacement, or double replacement.
4. recognize a chemical change in a laboratory setting.
5. apply the Law of Conservation of Mass to an experiment involving a chemical change.
6. balance equations by inspection.
7. complete simple chemical equations of the types listed earlier in this module.
8. convert a word equation into a symbolic equation.
9. draw "pictures in the mind" to represent the types of chemical reactions listed in this module.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Types of Reactions

Introduction

A large variety of chemical reactions goes on around us at all times. Classifying these reactions helps to make sense of them and enables us to predict the products. This activity gives an opportunity to learn and apply one classification scheme. There are other classification schemes that you will learn as you continue your study of chemistry.

Purpose

The purpose of this laboratory activity is to discover the characteristics of different types of reactions.

Safety

1. Wear eye protection and aprons throughout the activity.
2. Hydrochloric acid is caustic and corrosive. If you spill any on your skin wash it off with water and notify the teacher. Skin contact with other chemicals and solutions should be avoided.
3. Dispose of the chemicals as your teacher directs.

Procedure

Before beginning the laboratory work, read through the entire procedure and prepare a data table for each part of the activity. Record all data and observations in the table.

Part I

1. Determine the mass of a watch glass together with a piece of compacted copper wool approximately the diameter of a quarter.
2. Hold your tongs in the hot part of a burner flame for about one minute to remove any contamination.
3. With the tongs hold the copper in the hot part of the flame for 5 min. Make observations during the heating as well as after.
4. Air cool the copper for one minute before you return it to the watch glass.
5. When the watch glass and copper are cool determine the mass.
6. Dispose of the copper as your teacher directs.
7. Repeat the entire experiment with a sample of steel wool.

Part II

1. Add copper(II) sulfate pentahydrate, $CuSO_4 \cdot 5H_2O(s)$, to a depth of about 1 cm to a clean dry test-tube (13- x 100-mm). Observe the appearance of the copper(II) sulfate pentahydrate and record.
2. Determine the mass of the test-tube and sample.
3. Tilt the test-tube, spreading the hydrate so it covers about half the length of the test-tube.

4. Use a test-tube holder to hold the test-tube nearly horizontal. Move it back and forth through a cool flame (no inner blue cone) so that the entire sample is heated. Make observations while heating as well as after you are finished. (You have heated too much if the solid begins to darken.) Heating will be completed after about 2 min.
5. Allow the test-tube and contents to cool and determine its mass.
6. Use your dropper to add a drop of water to the substance in the test-tube. Record your observations.
7. Dispose of the chemicals as your teacher directs.
8. Repeat this part of the activity with calcium chloride hexahydrate, $CaCl_2 \cdot 6H_2O(s)$ or ammonium carbonate, $(NH_4)_2CO_3(s)$.

Part III

1. Place 5 mL copper(II) sulfate solution, $CuSO_4(aq)$, in a large test-tube (16 x 150 mm). Add a strip of magnesium, Mg. Make observations immediately and after 5 min.
2. Place 5 mL hydrochloric acid solution, HCl(aq), in a large test-tube. Add one piece of zinc, Zn. Make observations immediately and after 5 min.
3. Dispose of the chemicals as your teacher directs.

Part IV

1. Pour 3 mL barium chloride solution, $BaCl_2(aq)$, in a test-tube (13- x 100-mm). Add 3 mL sodium sulfate solution, $Na_2SO_4(aq)$.
2. Pour 3 mL cobalt(II) nitrate solution, $Co(NO_3)_2(aq)$, in a test-tube (13- x 100-mm). Add 3 mL sodium phosphate solution, $Na_3PO_4(aq)$.
3. Dispose of the chemicals as your teacher directs.
4. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

Part I

1. What evidence do you have for chemical change in each reaction?
2. Account for the change in mass in each reaction.
3. What characteristics do the reactions have in common?
4. Write a word equation to summarize what happened in each of the chemical changes.
5. Write a symbol equation for the reaction that took place. Look up the formulas if necessary.
6. This type of reaction is sometimes called *synthesis* or *combination*. Explain why this is an appropriate name.

Part II

Answer Questions 1-5 (from *Part I*) for these two reactions.

6. This type of reaction is called *decomposition*. Explain why this is an appropriate name.

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Part III

Answer Questions 1 and 3-5 (from *Part I*) for these reactions.

6. This type of reaction is called *single replacement*. Again, explain why this is a good name.

Part IV

Answer Questions 1 and 3-5 (from *Part I*) for these reactions.

6. This type of reaction is called *double replacement*. Explain why this is a good name.

Implications and Applications

1. Classify the following reactions as single replacement, double replacement, synthesis or decomposition.
   a. $NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
   b. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$
   c. $2HgO(s) \rightarrow 2Hg(l) + O_2(g)$
   d. $2Al(s) + 3CuSO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3Cu(s)$
2. Is it generally possible to classify a chemical reaction from observable evidence? Explain.
3. What other information would be helpful in classifying reactions?

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Types of Reactions

Level

General or basic student

Major Chemical Concept

Many chemical reactions can be classified into four types—Synthesis, Decomposition, Single Replacement, Double Replacement.

Expected Student Background

Students should be able to use a balance and burner correctly.

Time

45 min

Safety

See *Student Version*

Materials (For 24 students working in pairs)

Nonconsumables

- 24 Test-tubes, 16- x 150-mm
- 48 Test-tubes, 13- x 100-mm, or twelve 24-well spot plates
- 12 Burners
- 12 Tongs
- 12 Test-tube holders
- 12 Medicine droppers

Consumables

- Copper wool, 12 g
- Steel wool, 12 g
- Zinc metal, 6 g
- Magnesium ribbon, 0.6 g (36 cm)
- Copper sulfate, $CuSO_4 \cdot 5H_2O$, 18 g
- Calcium chloride, $CaCl_2 \cdot 6H_2O$, 18 g
- Ammonium carbonate $(NH_4)_2CO_3$, 18 g
- 3 M Hydrochloric acid, HCl, 60 mL (25 mL conc. HCl diluted to 100 mL)
- 0.1 M Copper sulfate, $CuSO_4$, 60 mL (2.5 g $CuSO_4 \cdot 5H_2O$ diluted to 100 mL)
- 0.1 M Barium chloride, $BaCl_2$, 36 mL (2.4 g $BaCl_2 \cdot 2H_2O$ diluted to 100 mL)
- 0.1 M Sodium sulfate, Na_2SO_4, 36 mL (3.2 g $Na_2SO_4 \cdot 10H_2O$ diluted to 100 mL)
- 0.1 M Cobalt nitrate, $Co(NO_3)_2$, 36 mL (2.9 g $Co(NO_3)_2 \cdot 6H_2O$ diluted to 100 mL)
- 0.1 M Sodium phosphate, Na_3PO_4, 36 mL (3.8 g $Na_3PO_4 \cdot 12H_2O$ diluted to 100 mL)

Advance Preparation

30 min to prepare solutions. Concentrations are not critical. Other concentrations may be substituted if available.

Pre-Laboratory Discussion

Review the evidence for chemical change. You may wish to demonstrate reactions that exhibit the types of evidence that students can use to decide whether a chemical reaction has occurred. Project Seraphim "Periodic Table" videodisc or ACS "Closeup on Chemistry" videodisc present excellent reactions.

Teacher-Student Interaction

As students are doing the activity, circulate through the laboratory and make sure they are attentive to all changes. Students may focus on one type of change (sparks given off when copper burns) and entirely miss another type of change (the copper turned black).

Anticipated Student Results

Part I

The steel wool shows a small increase in mass. The copper wool often does not because small fragments of copper escape.

Part II

Both hydrates lose mass. Both show steam and condensation. The copper(II) sulfate pentahydrate changes from blue to white. There is no color change with the calcium chloride hexahydrate.

Part III

The magnesium metal is replaced by a "coppery" colored solid and the blue color of the solution fades. Hydrogen bubbles may be noticed on the surface of the magnesium at the beginning of the activity. They are formed because the pH of 0.10 M $CuSO_4$ is about 4. Bubbles are formed on the surface of the zinc. If the piece of zinc is small enough it may disappear.

Part IV

A precipitate is immediately formed when each pair of solutions is mixed.

Answers to Data Analysis and Concept Development

Part I

1. The color changed.
2. The mass increased, so some substance must have been added.
3. In both reactions the substance changed color and increased in mass.
4. Copper + oxygen → Copper oxide*
   Iron + oxygen → Iron oxide
   *This would be a good time to teach students the use of Roman numerals to identify the oxidation states of metals.
5. See equations in *Post-Laboratory Discussion*.
6. Two elements are combined to form a single substance or a single substance is synthesized from the elements. *NOTE: Although these reactions started with elements, synthesis or combination also applies to using two or more compounds to synthesize another compound.*

Part II

1. Color change, water formed
2. The water was removed.
3. Water was formed in each reaction.
4. Copper(II) sulfate pentahydrate → Copper(II) sulfate + Water
   Calcium chloride hexahydrate → Calcium chloride + Water

5. See equations in *Post-Laboratory Discussion*.
6. Two substances were made from a single substance.

Part III

1. In one reaction a color change occurred and a new solid was formed. In the other, gas bubbles appeared.
2. Not applicable
3. One substance disappeared and another with different properties appeared.
4. Copper(II) sulfate + Magnesium → Copper + Magnesium sulfate
   Hydrochloric acid + Zinc → Hydrogen + Zinc chloride
5. See equations in *Post-Laboratory Discussion*.
6. When looking at the equation one can see that one element replaced another in a compound.

Part IV

1. A precipitate was formed.
2. Not applicable
3. Solutions of two compounds were mixed and a precipitate was formed.
4. Barium chloride + Sodium sulfate → Barium sulfate + Sodium chloride
   Cobalt(II) nitrate + Sodium phosphate → Sodium nitrate + Cobalt(II) phosphate
5. See equations in *Post-Laboratory Discussion*.
6. Looking at the equations and formulas one can see that the positive ions were exchanged. One of the products was insoluble and precipitated.

Answers to Implications and Applications

1. a. Double replacement
   b. Synthesis
   c. Decomposition
   d. Single replacement
2. Most often it is not. One can only observe whether a chemical change has occurred.
3. Names and formulas of starting materials would be very helpful.

Post-Laboratory Discussion

Part I

$$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$$

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

Use logic to show how a decrease in mass or no change in mass would not support the conclusion that a chemical reaction has occurred.

Example: If mass stays the same then the only possibility is copper (reactant) → copper (product), implying no chemical change.

Example: If mass decreases then the only possibility is that copper (reactant) → something less than copper (product); logically that is not possible because copper is an element and is not decomposed in normal chemical reactions. Emphasize the observations of a new product being formed on the surface of the copper wool.

Part II

$$\begin{aligned} CuSO_4 \cdot 5H_2O(s) &\rightarrow CuSO_4(s) + 5H_2O(l) \\ CaCl_2 \cdot 6H_2O(s) &\rightarrow CaCl_2(s) + 6H_2O(l) \\ (NH_4)_2CO_3(s) &\rightarrow 2NH_3(g) + CO_2(g) + H_2O(l) \end{aligned}$$

The apparent loss of mass needs to be examined since it would contradict the Law of Conservation of Mass. Make a drawing of the copper(II) sulfate (simple is fine) and use the color change from bright blue to near-white as evidence for change. The "loss" of water molecules can be verified by the steam and condensation seen at the mouth of the test-tube while heating. Calcium chloride will not show a color change, but will show condensation as evidence for change.

Part III

$$\begin{aligned} CuSO_4(aq) + Mg(s) &\rightarrow Cu(s) + MgSO_4(aq) \\ Zn(s) + 2HCl(aq) &\rightarrow H_2(g) + ZnCl_2(aq) \end{aligned}$$

The loss of the blue color of the copper(II) sulfate solution shows that copper ions have changed to copper metal. If students use only 1.0 cm of magnesium metal this will be more evident. Be sure that the students understand that all materials are present originally and are not created in the laboratory activity. Have them predict the identity of products that form. It should be noted that a "single replacement" reaction is a redox reaction (see *Oxidation-Reduction* module).

Part IV

$$\begin{aligned} BaCl_2(aq) + Na_2SO_4(aq) &\rightarrow BaSO_4(s) + 2NaCl(aq) \\ 3Co(NO_3)_2(aq) + 2Na_3PO_4(aq) &\rightarrow Co_3(PO_4)_2(s) + 6NaNO_3(aq) \\ H_2SO_4(aq) + NaOH(aq) &\rightarrow NaHSO_4(aq) + H_2O(l) \end{aligned}$$

Use these reactions to introduce exchange reactions. Students can practice predicting products. Emphasize using correct formulas. It should be noted that a "double replacement" reaction involves precipitation reaction and does not involve nonreacting spectator ions.

Possible Extensions

These reactions can be used again when students are introduced to moles and quantitative results of reactions. For example, water of hydration laboratory activities can be done quantitatively and total moles of water of hydration can be calculated.

Assessing Laboratory Learning

These reactions can be demonstrated and explanations given verbally. Formulas for the reactants may be helpful. You may wish to choose reactions from those given in the activity or demonstrations section or others from your repertoire.

1. Identify the type of reaction.
    a. A metal is heated in air. The product is a black powder. The mass of the powder is greater than the mass of the metal. *[synthesis]*
    b. Two clear colorless solutions are mixed. A precipitate forms immediately. *[double replacement]*
    c. A red powder is heated. A silvery liquid and a clear colorless gas forms. *[decomposition]*
    d. A metal and a clear liquid are mixed. A dark colored solid and a greenish solution result. *[single replacement]*
2. Draw "pictures in the mind" to illustrate each reaction studied in this laboratory activity.

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: Identifying Products of a Reaction

Introduction

The ability to classify chemical reactions helps us to predict the products of the reaction. However, only experimentation can confirm the identity of substances produced in a chemical reaction. In this activity two solutions are mixed to initiate a chemical reaction. The identity of the products is confirmed by flame tests and another chemical reaction.

Purpose

To observe the reaction of calcium chloride and potassium carbonate, predict the products, and confirm the identity of the products.

Safety

1. Wear eye protection throughout the laboratory activity.
2. Hydrochloric acid is caustic and corrosive. If you spill any on your skin wash it off with water and notify the teacher. Skin contact with other chemicals and solutions should be avoided.
3. Dispose of the chemical substances as your teacher directs.

Procedure

Before beginning the laboratory work, read through the entire procedure and prepare a data table. Record all data and observations in the table.

1. Weigh 2.5 g potassium carbonate, K_2CO_3, and 2 g calcium chloride, $CaCl_2$. Place in separate 100-mL beakers.
2. Add 25 mL distilled water to each beaker. Stir until all the solid is dissolved.
3. Number four 13 x 100-mm test-tubes from 1 to 4. Put 2 mL potassium carbonate solution in Tube 1 and 2 mL calcium chloride solution in Tube 2. Place a clean splint in Tubes 1 and 2.
4. Fold a piece of filter paper to fit the funnel.
5. Add the remaining solution of potassium carbonate to the remaining solution of calcium chloride. When the reaction is complete, filter the solution with pleated filter paper to hasten filtration. Remove the filtrate and wash the precipitate with 10 mL distilled water. If the filtrate is cloudy, filter again through the same filter paper.
6. Place 2 mL filtrate and a clean splint in Tube 3.
7. With a clean splint remove a small amount of precipitate from the filter paper. Put the splint and the solid in Tube 4. Add 2 mL distilled water.
8. Perform flame tests by holding each wet splint from Tubes 1-4 in the burner flame. Observe the first color you see. *Do not allow the splint to burn.* Dispose of the splints in the trash can.
9. Add three drops 3 M hydrochloric acid, HCl, to each solution in the test-tubes. Formation of bubbles indicates that the carbonate ion, CO_3^{2-}, was originally present.
10. If your teacher directs, set aside some of the filtrate to observe the crystals that form.

11. Dispose of the solutions and filter paper as your teacher directs.
12. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. What evidence do you have that a chemical change has occurred?
2. Describe the reactants and products.
3. Using the results of the test performed on the solutions, identify the precipitate formed and the substance found in the filtrate.

Implications and Applications

1. Draw pictures to represent:
    a. $CaCl_2$ before the water is added.
    b. $CaCl_2$ after the water is added.
    c. K_2CO_3 before the water is added.
    d. K_2CO_3 after the water is added.
    e. The substance in the filter paper.
    f. The substance dissolved in the filtrate.
2. Lead(II) chromate, $PbCrO_4$, can be used as a pigment to give a yellow color to paint. What compound could be added to $Pb(NO_3)_2$ to produce this pigment?
3. The labels have come off two bottles of white solids in the chemistry storeroom. The labels are KCl and K_2CO_3. How could you determine which label belongs on which bottle?

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: Identifying Products of a Reaction

Level

General or basic student.

Major Chemical Concept

The products of a double replacement reaction can be determined experimentally. The amounts of reactants are equimolar.

Expected Student Background

Students should be familiar with filtration and be able to use a burner safely.

Time

45 min

Safety

See *Student Version.*

Materials (For 24 students working in pairs)

Nonconsumables

24 Beakers, 100-mL
12 Funnels
12 Erlenmeyer flasks, 125-mL
48 Test-tubes, 13- x 100-mm
12 Burners
12 Stirring rods
12 Graduated cylinders, 25-mL
12 Blue cobalt glass squares, if available, for viewing the potassium flame test.

Consumables

Calcium chloride, $CaCl_2$, 24 g
Potassium carbonate, K_2CO_3, 30 g
Filter paper
3 M Hydrochloric acid, HCl, 12 mL (25 mL conc. HCl diluted to 100 mL)
48 Splints

Advance Preparation

20 min to assemble the materials.

Pre-Laboratory Discussion

Review the evidence for chemical change and double replacement reactions. Emphasize that the only way we can be sure about the identity of products is to test them. Discuss flame tests if students are unfamiliar with them. The only use of the flame test in this activity is to identify potassium and calcium ions.

Teacher-Student Interaction

During the activity focus students' attention on the difference in properties of the reactants and products. Caution them to observe the flame test before the splint catches on fire. They should compare the colors of the flame tests on the known starting materials and the products.

Anticipated Student Results

Sample Data Table

	Test-tube Number			
Test	**1**	**2**	**3**	**4**
Flame	Violet	Red	Violet	Orange
Acid	Bubbles	No rxn	No rxn	Bubbles

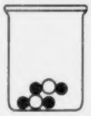
Answers: Data Analysis and Concept Development

1. A precipitate is formed.
2. Observations of reactants:

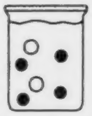
   Potassium carbonate - white solid
   Calcium chloride - white solid
   Both dissolved in water
   Observations of products
   One substance precipitates and the other dissolves in water, passing through the filter paper.
3. The precipitate is calcium carbonate; the substance in the filtrate is potassium chloride.

Answers: Implications and Applications

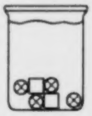
1a.



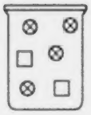
1b.



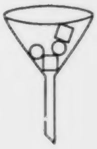
1c.



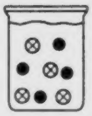
1d.



1e.



1f.



Ca^{2+} = ○   K^{+} = ⊗
Cl^{-} = ●   CO$_3^{2-}$ = □

2. $K_2CrO_4(aq)$ or $Na_2CrO_4(aq)$
3. Make a solution of a small portion of each solid. Test each solution with a solution of calcium chloride. The one that precipitates is K_2CO_3. Alternatively, add dilute hydrochloric acid to both. The one that forms bubbles is K_2CO_3.

Post-Laboratory Discussion

The discussion should focus on student answers to *Data Analysis and Concept Development* questions. The Pictures in the Mind section is particularly important in helping students develop the concept of the particulate nature of matter. These pictures can also be used to explain anomalous results arising from incomplete separations.

Possible Extensions

1. Students may plan an investigation to determine the identity of products produced in one of the demonstrations that follow this section.
2. The use of flame tests as an identification tool may be studied further.

Assessing Laboratory Learning

1. Extension 1 could serve as an evaluation. Be sure to use reactions involving chemistry that students can reasonably be expected to know.
2. Pictures in the Mind of different systems (Example: Demonstration 3) can be used to evaluate whether students have a reasonable concept of what occurs at the molecular level.

DEMONSTRATIONS

Purpose

These demonstrations allow students to observe a wide variety of chemical reactions, most of which should not be performed by students for safety reasons. Be sure to keep students actively involved by directing them to make observations and consider questions during demonstrations. They could be asked to write down observations.

Safety

Wear eye protection and an apron throughout the demonstrations. Thoroughly wash your hands after concluding the demonstrations. Additional safety precautions are given for each demonstration.

Demonstration 1: Magnesium Burning

Materials

Magnesium ribbon, Mg, 3-5 cm strip
Burner
Crucible tongs

Safety

Caution students not to look at flame directly.

Directions

Light the burner. While holding the magnesium with the crucible tongs, light the magnesium in the burner flame. Point out the evidence that a reaction has occurred.

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Demonstration 2: A Penny in Nitric Acid

Materials

Erlenmeyer flask, 1000-mL
Penny, pre-1983
Conc. Nitric acid, HNO_3, 30 mL

Safety

Conduct the reaction in the hood—the fumes are toxic. Nitric acid is highly corrosive.

Directions

This reaction must be performed in a fume hood. (The fumes of nitrogen oxides are toxic.) With the hood operating, place the penny in the beaker. Add nitric acid. Observe all color changes and products released. After the penny has reacted, add a small quantity of water to the beaker to bring about a color change. (In this reaction colorless NO is initially formed; it subsequently reacts with O_2 to form colored NO_2, which is what students observe.) They also note the color of the solution.

$$Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

Demonstration 3: Formation of a Precipitate

Materials

Beaker, 250-mL
0.1 M Lead nitrate, $Pb(NO_3)_2$, 50 mL [Dissolve 1.7 g $Pb(NO_3)_2$ in distilled H_2O and dilute to 50 mL]
0.1 M Potassium iodide, KI, 50 mL [Dissolve 1 g KI in distilled H_2O and dilute to 50 mL]

Safety

Properly dispose of the lead compound formed.

Directions

Simultaneously pour the two solutions into the beaker. The brilliant yellow precipitate lead iodide, PbI_2, forms. Students should recognize that a chemical reaction has occurred.

$$Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$$

Demonstration 4: Replacement Reaction

Materials

Copper wire, Cu, 5 cm strip
Large test-tube, 15- x 120-mm
0.1 M Silver nitrate, $AgNO_3$, 20 mL [Dissolve 0.3 g $AgNO_3$ in distilled H_2O and dilute to 20 mL]

Safety

Properly dispose of the products.

Directions

Pour silver nitrate solution into the large test-tube. Support the test-tube in a rack. Drop in the copper wire. The reaction will start immediately, but will be more obvious after a few minutes as the silver begins to deposit on the surface of the copper wire. The observed formation of the new product is evidence of reaction.

$$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$

Demonstration 5: Sodium with Water

Materials

Petri dish
Water
Sodium metal, Na, no larger than a grain of uncooked rice—see *Safety*

Safety

This reaction is striking and demonstrates vividly to students that chemical reactions not only result in a change in properties, but also often release energy. However, be aware that *without the precautions specified here, this demonstration can be dangerous.*

Directions

Fill the Petri dish about half full with cold water. Be sure that the water is *cold* and the piece of sodium is no larger than a grain of uncooked rice. Drop the piece of sodium into the water. Observe the reaction. Sometimes the generated hydrogen will actually burn. Small scrapings or cuttings of sodium may be disposed of by dropping them in alcohol. Some chemical companies that supply schools sell an alloy of approximately 10% sodium and 90% lead that can be used in place of pure sodium.

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

Demonstration 6: Endothermic Reaction

Materials

Barium hydroxide, $Ba(OH)_2$, 10 g
Ammonium thiocyanate, NH_4SCN, 10 g
Erlenmeyer flask, 250-mL

Safety

Properly dispose of the products.

Directions

Place the barium hydroxide in the flask. [*CAUTION: $Ba(OH)_2$ is* very *toxic.*] Add the ammonium thiocyanate. Swirl the two solids for about 20-30 sec. Give the flask to a student to hold. Evidence for the reaction is an unexpectedly cold container, indicating an endothermic process.

$$Ba(OH)_2(s) + 2NH_4SCN(s) \rightarrow 2NH_3(g) + H_2O(l) + Ba(SCN)_2(s)$$

Demonstration 7: Floating Pennies

Materials

Pennies minted after 1983
6 M Hydrochloric acid, HCl
Large graduated cylinder

Safety

Hydrochloric acid is a corrosive substance.

Directions

Pour hydrochloric acid into the cylinder so it is about 3/4 full. Use a triangular file to scrape the edge of each penny to expose the zinc core. It seems to work best to scrape twice—on opposite edges of the penny. Drop the pennies into the hydrochloric acid. Bubbles will be seen immediately, but the pennies exhibit unusual behavior after about 24 hours. They will float and sink as their density approaches that of the surrounding acid. The hollow pennies are also interesting to examine after the reaction is completed. You can prepare "day old" pennies for this purpose.

$$Zn(s) + 2HCl(aq) \rightarrow H_2(g) + ZnCl_2(aq)$$

Demonstration 8: Indicators

Materials

Various indicator solutions such as bromthymol blue, methyl orange, bromcresol green, *etc.*
0.1 M solutions of HCl and NaOH
Small beakers or small Petri dishes, same number as number of indicators used

Directions

Place beakers on an overhead projector. Add just enough acid solution to cover the bottom of each beaker. Add an indicator to each. Note color. Then add base dropwise until color change occurs. Be sure to swirl contents as base is added.

Demonstration 9: Aladdin's Lamp

Materials

Large round-bottomed flask or 500-mL Erlenmeyer flask
30 % Hydrogen peroxide, H_2O_2, 50 mL *(see* Cautions *below)*
Potassium iodide, KI, 2 g

Safety

30% hydrogen peroxide is dangerous. It should be very carefully handled and used. Wear gloves when handling. Under no circumstances should students handle 30% hydrogen peroxide.

Directions

Support the flask with a clamp on a ring stand. Pour the hydrogen peroxide into the flask. Add the KI. Within a few seconds, the KI generates so much heat as it reacts with the hydrogen peroxide that it boils the water! Using a fancy spouted kettle might be amusing.

Discussion

This decomposition takes place in two steps:

$$H_2O_2(aq) + I^-(aq) \rightarrow 2H_2O(l) + IO^-(aq)$$
$$H_2O_2(aq) + IO^-(aq) \rightarrow H_2O(l) + O_2(g) + I^-(aq)$$

The reaction is exothermic and produces oxygen gas and steam.

Demonstration 10: Ions Need to Get Together

Materials

2 Erlenmeyer flasks, 125-mL
Erlenmeyer flask, 250-mL
Potassium iodide, KI, 2 g
Copper(II) sulfate pentahydrate, $CuSO_4 \cdot 5\ H_2O$, 2 g

Directions

Add the blue solid, $CuSO_4 \cdot 5H_2O$, and the white solid KI to a flask. Shake them together; let students see that nothing is happening. The two salts are still as they were originally. Put the mixture of two salts in a mortar and grind vigorously with a pestle. A brown color should be visible. Point out the effort it took to get the substances to react. Then dissolve 1 to 2 g of each salt in separate 100 mL samples of water in 125-mL flasks. Pour the two solutions together. Students observe an immediate, obvious formation of a brown "gunk." This demonstrates that ions need to get together to react. In the solid state they could not do that, but in solution the ions could readily "find" each other.

$$4KI(aq) + 2CuSO_4(aq) \rightarrow 2K_2SO_4(aq) + 2CuI(s) + I_2(s)$$

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. What evidence can be observed to indicate that a chemical change has occurred? *[Color change, change in physical state, energy absorbed or released.]*
2. Why is it useful to be able to classify chemical reactions? *[Classification enables us to organize information and ultimately to predict the results of similar types of changes.]*
3. Is there only one type of classification that might be used for chemical reactions? *[No, there are many different ways to classify reactions. Many of these will be used as knowledge of chemistry increases. It is always necessary to state the criteria used for the classification.]*
4. What future classification schemes are suggested by this study? *[Is energy absorbed or released? How fast is the reaction? What kind of products are produced? What kind of process is involved in the change?]*

Pictures in the Mind

Which pictures in Figure 1 represent:

a. Elements
b. Solids
c. Gases
d. Compounds
e. Precipitates
f. Gaseous elements
g. Solid elements
h. Liquid elements

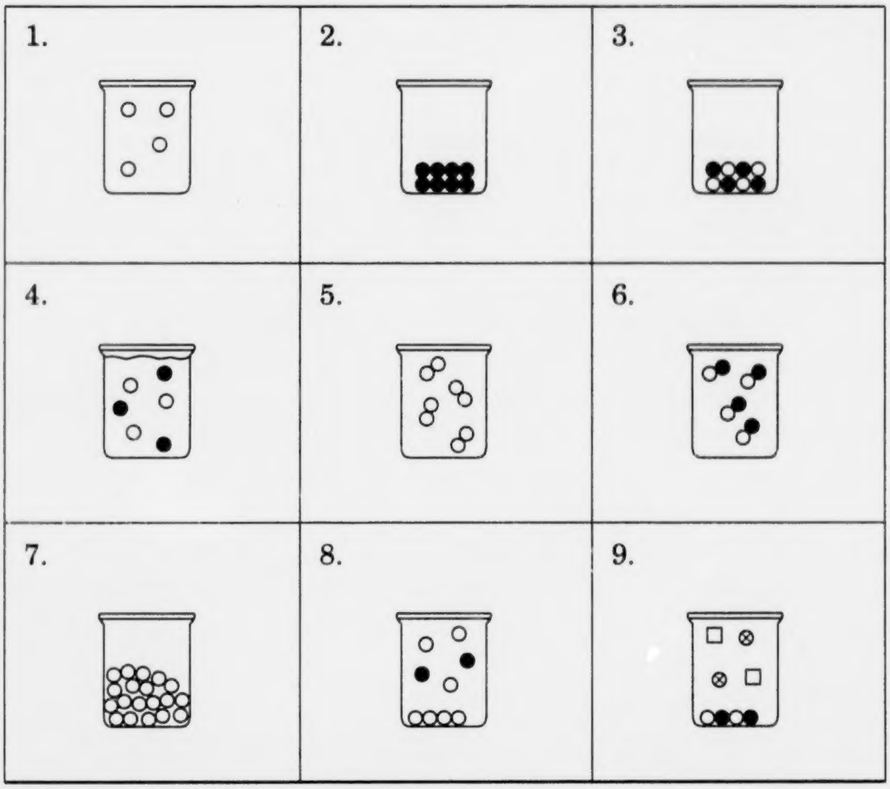


Figure 1. Representation of "Pictures in the Mind." *These molecular conglomerates can be made from Perler's "Big Beads for Little Fingers" (available at craft shops) and suspended in a beaker by thread from cardboard or a heavy paper lid.*

Student Activity

Formulas, Equations, Conservation of Mass

Materials

Zip-closure sandwich-size plastic bags

Paper clips—jumbo, regular, small, assorted colors (at least four different types are needed); use at least six clips of each kind.

Procedure

1. Have students work in groups of two or three. Instructions can be both verbal and written on the board, as the activity progresses. Give bag of clips to each group.
2. Use symbols J for jumbo, R for regular, Sm for small, and appropriate symbols for colored clips—Bl for blue; Wh for white; Rd for red (or any other designations you find satisfactory).
3. Give directions for the activity by having one student in each group make a particular model. Then instruct the other student to look at it to see if it is what would be expected. If not, have them agree on a structure and explain verbally to each other why it is like it is. The class can give feedback when groups do not agree.
4. Start with instructions to show an atom of an element—R, for example.
5. Proceed to more complex examples, *e.g.*, a diatomic molecule J_2, multiples of an element, 4 J, and then work into compounds, RBl, JWh_2, $J_2(BlRd)_3$.
6. Invite students to carry out chemical reactions involving a 1:1 mole ratio:

$$\begin{aligned} R + Bl &\rightarrow RBl \\ R + JWh &\rightarrow J + RWh \\ JBl + RWh &\rightarrow JWh + RBl \end{aligned}$$

7. To show the necessity for conservation of mass, proceed to:

$$\begin{aligned} 2R + Wh_2 &\rightarrow 2RWh \\ 2J + 2RRd &\rightarrow R_2 + 2JRd \end{aligned}$$

TIPS FOR THE TEACHER

Language of Chemistry

chemical reaction process by which one or more substances change into new substances.

compound substance composed of two or more elements.

decomposition reaction reaction that results in two or more substances being formed from a single substance.

double replacement reaction reaction in which the positive ions of two compounds are exchanged. "Displacement" is a synonym for "replacement."

element substance that cannot be decomposed into a simpler substance.

mixture two or more substances that are not chemically combined and does not have a fixed set of properties.

precipitate solid formed in a chemical reaction involving solutions.

single replacement reaction reaction in which one element replaces another.

solution homogeneous mixture.

substance sample of matter that has a uniform set of properties and a definite composition. Examples of substances are elements and compounds.

synthesis reaction two or more substances form a single substance in a reaction. Synonyms are "composition" and "combination."

Pattern Recognition

Teaching students to classify reactions is an exercise in pattern recognition.

Synthesis (combination; composition)

Element + Element → Compound

1. Iron + Sulfur → Iron(II) sulfide
2. Mercury + Oxygen → Mercury(II) oxide
3. Phosphorus + Oxygen → Diphosphorus pentoxide
4. Potassium + Chlorine → Potassium chloride

Compound + Compound → More complex compound

1. Magnesium oxide + Carbon dioxide → Magnesium carbonate
2. Sulfur dioxide + Water → Sulfurous acid

Decomposition

Compound → Two or more simpler substances

1. Water → Hydrogen + Oxygen
2. Calcium carbonate → Calcium oxide + Carbon dioxide
3. Sodium nitrate → Sodium nitrite + Oxygen
4. Copper(II) sulfate pentahydrate → Copper(II) sulfate + Water

Single replacement

Element + Compound → Element + Compound

1. Iron(III) oxide + Carbon → Iron + Carbon monoxide
2. Aluminum + Sulfuric acid → Aluminum sulfate + Hydrogen
3. Copper(II) sulfate + Iron → Copper + Iron(II) sulfate
4. Sodium + Water → Sodium hydroxide + Hydrogen

Double replacement

Two compounds → Two new compounds

1. Sodium carbonate + Calcium hydroxide → Sodium hydroxide + Calcium carbonate
2. Iron(II) sulfide + Hydrochloric acid → Hydrogen sulfide + Iron(II) chloride
3. Magnesium hydroxide + Sulfuric acid → Water + Magnesium sulfate
4. Copper(II) nitrate + Sodium hydroxide → Copper(II) hydroxide + Sodium nitrate

Common Student Misconceptions

1. **"Law of Conservation of Mass does not apply to atoms."**

   Students may be able to repeat the Law of Conservation of Mass, but see no problem with atoms disappearing or appearing to balance equations.

2. **"Elements can form other elements."**

   Several students when questioned about the appearance of copper on magnesium ribbon in copper sulfate remarked that the brown solid came from the magnesium. They thought it was somehow present in the magnesium even though they agreed that magnesium was an element.

3. **"All solutions are pure liquids."**

   Most students, early in the year, do not distinguish between a solution and a pure liquid. Teachers and textbooks perpetuate this notion by labelling solutions as 0.1 M $CuSO_4$ before students understand anything about molarity. It would probably be better to label such a solution as "$CuSO_4$ solution" and to demonstrate for students how such solutions are prepared. One could also demonstrate that when a pure liquid is evaporated there is no residue, but when a solution is evaporated there is often a residue.

4. **"There is such a thing as an unbalanced chemical equation."**

   An equation, by definition, is an expression in which the components on either side of the equality sign are identical. For example, in the equation 4x = 32, we are told that multiplication of x by 4 yields the quantity 32. It follows that x must be 8, and that the equation can be rewritten as 4 x 8 = 32.

   A chemical equation is not a mathematical equation in the strict sense. The left side of the equation represents the reactants and the right side the products. They are joined by an arrow or "yields" sign that, for all intents and purposes, may be taken as an equality sign. The chemical equation is an expression of the Law of Conservation of Matter. Equal numbers of atomic representations must appear on both sides of the arrow. In the expressions given below, one is a chemical equation and the other is a chemical statement:

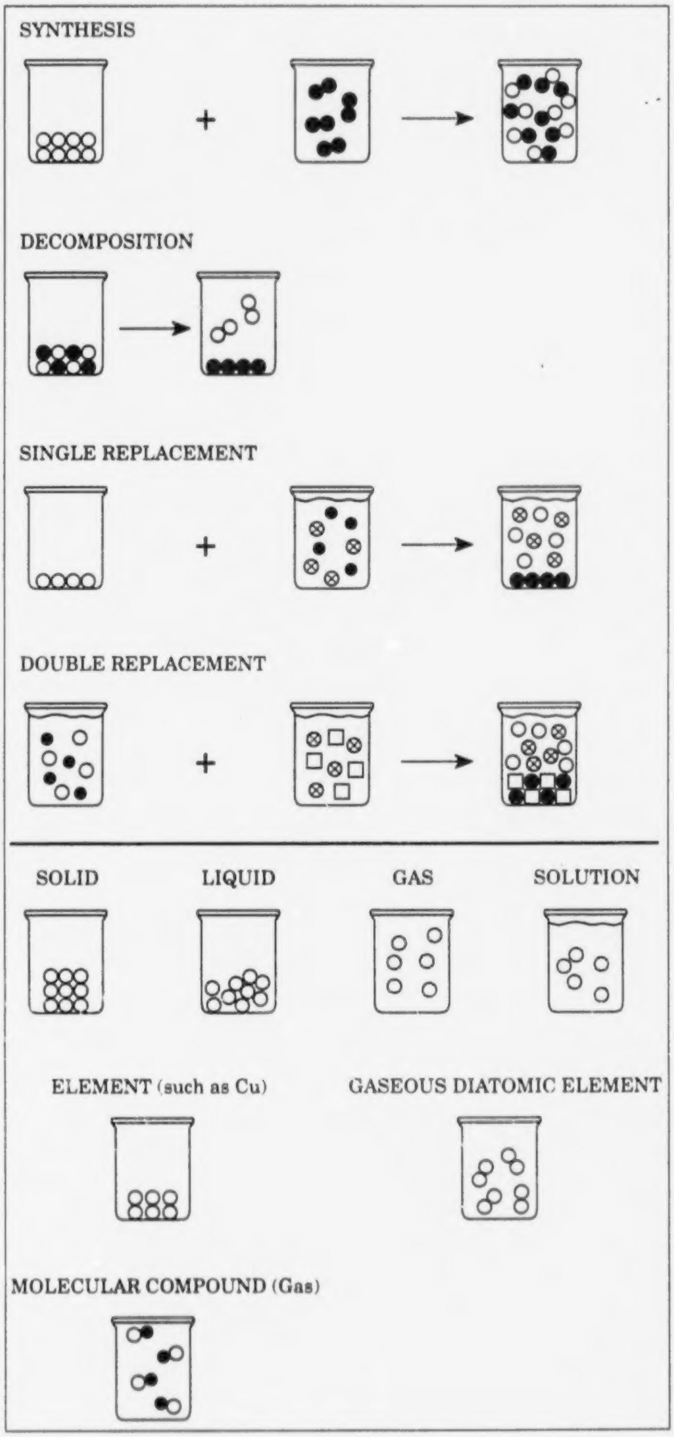
   a. $CaCO_3(s) + HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

   b. $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

   In expression (a), the number of chlorine atoms and hydrogen atoms are unequal on both sides of the arrow. Placing a 2 in front of the HCl in the expression achieves the desired (and required) equality. Hence, expression (b) fulfills the definition of an equation, but expression (a) does not. Expression (b) is often called a "balanced chemical equation," and expression (a) is called an "unbalanced chemical equation." Strictly speaking, expression (a) is not an equation at all since the equality requirement is not fulfilled, and expression (b) is an equation and therefore does not need the adjective "balanced" since it is a redundancy, by definition. Common usage allows us to call both expressions equations, but according to our definition, there is no such thing as an unbalanced equation.

Pictures in the Mind

"Pictures in the mind" is a useful device to help students understand on a molecular level what is happening when a chemical reaction occurs. If attention is paid to the numbers of atoms that combine to form new substances and to balancing equations, it may help students later to attack the challenging (for many) subject of stoichiometry. Be sure to include containers for the pictures so students do not acquire incorrect ideas of solids and liquids. If you have not already introduced pictures for solids, liquids, gases, compounds, and elements, do so before reactions are considered.


SYNTHESIS
+
DECOMPOSITION
SINGLE REPLACEMENT
+
DOUBLE REPLACEMENT
+
SOLID
LIQUID
GAS
SOLUTION
ELEMENT (such as Cu)
GASEOUS DIATOMIC ELEMENT
MOLECULAR COMPOUND (Gas)

Solids:
Units touching and organized

Liquids:
Units touching but not organized

Gases:
Space between units (molecules)

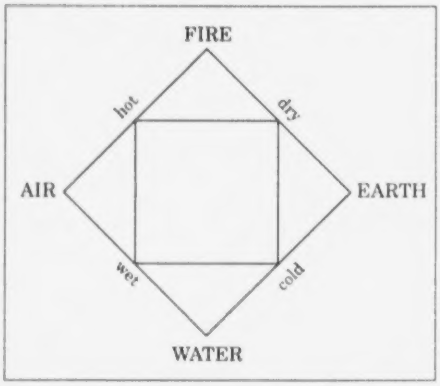
Solutions:
Units separated by solvent molecules

Figure 2. Pattern recognition for "Pictures in the Mind."

HISTORY: ON THE HUMAN SIDE

"The systems that confront the intelligence remain basically unchanged through the ages, although they assume different forms...There is nothing so disastrous in science as the arrogant dogmatism that despises the past and admires nothing but the present." Hoffer.

Chemistry has its beginnings in the early civilizations of Egypt, India, and China. However, the basis for the developments that took place in European nations can be found in Aristotle's works from 350 B.C. In his *Theory of the four elements,* Aristotle connected earth, air, fire, and water with the physical qualities of cold, wet, hot, and dry.



Hovering behind these four elements was a shadowy and poorly defined fifth element, that would later be known as the 'philosopher's stone,' or *quinta essentia*, held to be the reason or driving power behind chemical reactions. The theory of four elements also tried to explain chemical reactions as the changing, or transmuting of one element into another. For example, water, the cold-wet element, can be transmuted into air through the application of heat. This theory of transmutation was soon extended to metals, leading to the centuries-long fruitless search for a method of converting common metals, such as lead, into gold.

The attempts to convert other metals or even organic substances, such as eggs, into gold are far too numerous to list. The actual attempts are documented in cryptic forms, since the alchemists guarded their research carefully and coded the findings, often in the form of secretive poetry:

> "Dissolve the Fixt, and make the Fixed fly,
> The Flying Fix, and then live happily."

Other methods for preparing gold are described somewhat more directly. Many alchemists were successful at tricking and cheating their customers; the actual products of their chemistry were not gold, but gold-colored metals.

The concept of transmutation relied on a theory that all forms of matter are one in origin and have subsequently been transmuted. The alchemists were searching for methods to change the very form of matter. The elusive philosopher's stone can be compared to modern catalysts that stimulate chemical reactions without themselves being permanently changed. Despite the apparent failure to change common metals into gold, the alchemists, through research and experimentation on metals and organic substances, formed the foundation for the scientific revolution and the development of modern scientific theory in the 17th century.

However, despite the changes for which Galileo and Newton provided the impetus in the 17th century, chemistry remained rather unchanged. Despite the new scientific concepts of the world as a macrocosm, subject to describable laws of motion, and new studies of the living organism, the Aristotelian elements retained their hold on chemistry. Not until the chemical interactions of elements and their composition were better understood, could chemistry change drastically. The action of combustion and the nature of the so-called elements air and water were not accurately described until Lavoisier isolated oxygen and gained an understanding of combustion in the 1770s. The challenges faced by chemists to explain chemical phenomena from the 18th century to the present continues.

HUMOR: ON THE FUN SIDE

1. Message on a T-shirt: Looking for a reaction? Try me—I'm a chemist.

2. **THE H_2 OF AN O**

One thing I would like to do
Is draw a character sketch of CO_2...
Meet an angry LiH perhaps,
Or talk to a placid He gas...
Find an NaH that likes to play
And picnic on a sunny day,
Or an NO_3 that moaned and cried
When his kindly mother, HNO_3, died.
The world could hold
Satisfied Ne's
Capitalist Rb's
Socialist Pb's.

But then of course, on cloudy days
The H_2's would go off to war
And cruelly spread their acid horror
Among the lowly SO_4.

Yes, good is mixed with evil:
That is the way of life.
And if we're to have friendly O's
We must have H_2 strife.

Au bless us, every mole.

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3. **CHEMISTRY ROCK**

Chemistry, Chemistry, Chemistry Rock
We're so scientific, the work never stops
We balance equations around the clock
That's the Chemistry Rock!

Seventh bell time[1] is a swell time
To rock the lab away
Computations, calculations
Well, Mr. Evans[2] might give you an "A"

So listen here, lend an ear and we'll tell you why
It's chemistry at the top
It'll keep your mind working overtime
That's the Chemistry Rock!

[1] Put in your own time.

[2] Put in your teacher's name.

CHEM 13 NEWS, February 1982, p. 12

A Visit from Father Christmas

(A CHRISTMAS FABLE)

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Wilkes-Barre, Pennsylvania

'Twas the night before Xmas,
The lab was quite still;
Not a Bunsen was burning
(Nor had they the will).
The test-tubes were placed
In their racks with great care;
In hopes Father Chemistry
Soon would be there.

The students were sleeping
So sound in their dorms,
All dreaming of fluids
And Crystalline forms.
Lab-Aides in their aprons
And I in my smock
Were sitting, recov'ring
From semester-end shock.

When outside the lab
There arose such a roar
I leaped from my stool
And fell flat on the floor.
Out to the fire escape
All of us flew.
What was the commotion?
Not one of us knew.

The flood-lights shone out
O'er the campus so bright
It looked like old Stockholm
On Nobel Prize Night.
My fume-blinded eyes,
Then viewed (dare I say?)
Eight anions pulling
A water-trough sleigh.

And holding the bonds
Tied to each one of them
Was a figure I knew
As our own Papa Chem.
With speeds in excess
Of most X-rays they came.
As they Dopplered along
He called each one by name.

"Now Nitrite, now Phosphate,
Now Borate, now Chloride,
On Citrate, on Bromate,
On Sulfite and Oxide.
Forget what you know
Of that randomness stuff,
Let's go straight to that roof,
If you've quanta enough."

As fluids Bernoullian
Behave in a pinch
Those ions said "Alchemist,
This is a cinch,"
So up to the lab-roof
Those "chargers" they sped
With Pop Chemistry safe
In his water-trough sled.

Just a microsec later
Electroscopes showed
Charged particles coming
To our lab-abode.
We raced back inside,
And what do you think?
Down the fume-hood Pop Chem fell,
Right into the sink.

He was dressed in a lab-coat,
Quite ragged and old,
With removable buttons
(The style, we're told)
A tray-full of beakers
He clutched to his heart—
And under his arm
Was an orbital chart.

His eyes through his goggles
I just couldn't see
His hands were all yellow
From H-N-O-3
His head was quite bald
With a fringe all around
Like a ring test for iron,
That same shade of brown.

He puffed a cigar
With a smell not at all
Unlike the organic lab
Right down the hall
The smoke billowed forth
From his angular face
And with Brownian Movement
Enveloped the place.

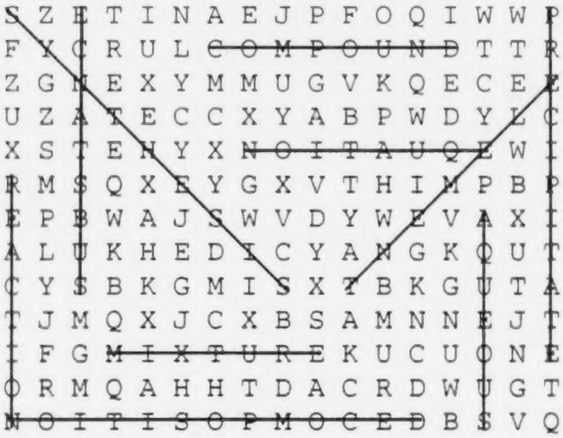
He was thin as a match
And not terribly tall
He wasn't the type
I'd expected at all
But a look at his clothes,
In the lab's harsh white light,
With their acid-burn holes—
He's a chemist all right.

He didn't say much
(He had no time to kill)
And filled all the test-tubes
With nary a spill.
Then placing them back
On the benches with care
He dashed to the fume-hood
And rose through the air.

He called to his team
And his ions took off
And kinetics took care
Of Pop Chem and his trough.
But I heard him cry out
As he flew down the street
"Merry Christmas to all!
May your stockrooms stay neat!"

CHEM 13 NEWS, December 1974, p. 2

5. Word Search (see *Appendix* for master copy)



S Z E T I N A E J P F O Q I W W P
F Y C R U L ~~C O M P O U N D~~ T T R
Z G N E X Y M M U G V K Q E C E E
U Z A T E C C X Y A B P W D Y L C
X S T E H Y X ~~N O I T A U Q E~~ W I
R M S Q X E Y G X V T H I M P B P
E P B W A J S W V D Y W E V A X I
A L U K H E D I C Y A N G K Q U T
C Y S B K G M I S X T B K G U T A
T J M Q X J C X B S A M N N E J T
I F G ~~M I X T U R E~~ K U C U O N E
O R M Q A H H T D A C R D W U G T
~~N O I T I S O P M O C E D~~ B S V Q

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Solid formed in a chemical reaction involving solutions.
2. Substance that cannot be decomposed into a simpler substance.
3. Reaction type in which two or more substances form a single substance in a reaction.
4. Reaction type in which two or more substances are formed from a single substance.
5. Process by which one or more substances change into new substances.
6. Two or more substances that are not chemically combined and do not have a fixed set of properties.
7. Substance composed of two or more elements.
8. Sample of matter that has a uniform set of properties and a definite composition.
9. Symbolism used to represent reactions that occur.
10. Adjective for a water solution.

Answers: 1. PRECIPITATE 2. ELEMENT 3. SYNTHESIS 4. DECOMPOSITION 5. REACTION 6. MIXTURE 7. COMPOUND 8. SUBSTANCE 9. EQUATION 10. AQUEOUS

6. See cartoons at end of module.

MEDIA

1. *Doing Chemistry* videodiscs are devoted to demonstrating chemical reactions. The wide variety and extensive list of chemical reactions make this an important teacher resource. "Reactions in a Bag," Experiment D45, provides an excellent introduction to the whole topic. Available from American Chemical Society, 1155 Sixteenth Street, N.W., Washington, DC 20036.
2. The *World of Chemistry* videotapes. World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational

Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224. A secondary school version of this series is available from WINGS for Learning/ SUNBURST, 101 Castleton Street, Pleasantville, NY 10570; (914) 747-3310; (800) 321-7511; (914) 747-4109 (FAX).

3. Software published by *JCE: Software*, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

    a. *KC? Discoverer with Knowledgeable Counselor,* by Daniel Cabrol, John W. Moore and Robert C. Rittenhouse. Special Issue 2, for IBM PS/2, PC compatible computers.

    b. *KC? Discoverer: Exploring the Properties of the Chemical Elements,* by Aw Feng and John W. Moore. Vol. I B, No. 1, for IBM PS/2, PC compatible computers.

    c. *KC? Discoverer?,* by Michael Liebl, Vol. IV A, No. 2, for all Apple II computers.

    d. *The Periodic Table Stack,* by Michael Farris. Vol. I C, No. 1, for the Apple Macintosh.

    e. *The Periodic Table* (Toolbook), by Paul F. Schatz, John C. Kotz and John W. Moore, in press. For Windows running on IBM PS/2 and PC-compatible computers.

    f. *REDOX: A Tutorial on REDuction /OXidation Equations,* by Derek Davenport, Paul Groves and Dale Jensen. Vol. III A, No. 1, for the Apple II computer.

    g. *PIRExS: Predicting Inorganic Reactivity Expert System,* by Jame P. Birk. Vol. III B, No. 1, for IBM PS/2 PC-compatible computers.

4. Videodiscs published by *JCE: Software*, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

    a. "Nitrogen Triiodide Decomposed" and "Thermite Reaction," two chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc II (double sided, 60 min.), Special Issue 4.

    b. "Similarities and Trends in Groups I – Noble Gases" and "Similarities and Trends in Groups II – Alkali Metals," "An Exothermic Reaction" and "An Endothermic Reaction," four chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc I (double sided, 60 min.), Special Issue 3.

    c. *The Periodic Table Videodisc* (single side, 30 min.), Special Issue 1.

    d. *Demonstrations in Organic Chemistry* (double sided, 60 min.). Special Issue 6.

5. *Redox*, by Helen Brooks and David Brooks. Distributed by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-2387 (voice) or (608) 262-0381 (FAX). Order Number VID 002.

6. *Closeup on Chemistry*, videodisc available from the American Chemical Society, 1155-16th Street, NW, Washington, DC 20036.
7. Concepts in General Chemistry Series, *Chemical Reactions* and *Reactions in Aqueous Solution*. Trinity Software, P.O. Box 960, Campton, NH 03223; (800) 352-1282.1.

Links/Connections

WITHIN CHEMISTRY

The mole concept was devised to explain (predict) the mass relationships involved in chemical reactions.

2. Whole fields of study in chemistry are based on a single class of reaction. Acid-base and redox are two prime examples (see these modules as well as the *Solubility and Precipitation* module).

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

1. The precipitate $Co_3(PO_4)_2$ prepared in *Laboratory Activity 1* is used as a ceramics coloring agent.
2. Photosynthesis (a process students may have encountered in biology) is an example of a combination or synthesis reaction. It is considerably more complex than reactions studied in this module, but the result of a series of reactions is that carbon dioxide and water combine to form glucose, a sugar.
3. Geology (see *Rocks, Minerals, and Gems* module)
4. Biochemistry (see *Enzymes* module)

TO THE CONTEMPORARY WORLD

1. The remnants of the space shuttle Challenger were under the sea for several weeks before being recovered. This caused the tapes carrying data and conversations of the crew to become covered with a foam-like cement of $Mg(OH)_2$. The tapes had been stored in a magnesium casing that reacted with seawater essentially as

   $$Mg(s) + 2H_2O(l) \rightarrow H_2(g) + Mg(OH)_2(s)$$

   catalyzed by salts in the water. The tapes were restored to a readable form by a painstaking removal of the cement layer using the double replacement reaction:

   $$Mg(OH)_2(s) + 2HNO_3(aq) \rightarrow 2H_2O(l) + Mg(NO_3)_2(aq)$$

2. Baking powder is added to doughs to generate the gas CO_2 inside the pastry to help it "fluff up." The gas comes from the reaction of sodium bicarbonate with an acid salt such as potassium acid tartrate.

   $$NaHCO_3(s) + KHC_4H_4O_6(s) \rightarrow NaKC_4H_4O_6(s) + H_2O(l) + CO_2(g)$$

3. Cooking an egg can also be viewed as a decomposition reaction, because the primary event is the breaking of chemical bonds. The white part of the egg is the protein, albumin. In its natural state, this very long chain of linked amino acids wraps around itself in a prescribed way. The prescribed wrapping is held in place by chemical bonds between adjacent parts of the chain. Heating breaks these "cross-linking" bonds, the molecular chain unwraps, stretches out, and then rebonds randomly to other unwrapped protein molecules. This whole process is seen as the clear, slimy liquid turning into a white pliable solid. Biochemists call it denaturation.

4. Rusting of iron and steel is a chemical reaction readily recognized by everyone. Overall, it can be classified as a synthesis reaction.

   $$4Fe(s) + 3O_2(g) + 6H_2O(l) \rightarrow 2Fe_2O_3{\cdot}3H_2O(s)$$

   The product, rust, is a hydrated form of iron(III) oxide. The actual sequence of reactions that yields this overall change is quite complex, and not yet completely understood. An interesting characteristic is that the process is catalyzed by salt. One possible explanation of most of the data is as follows.

   The metallic iron in contact with water is oxidized

   $$Fe \rightarrow Fe^{2+} + 2\,e^-$$

   These electrons then reduce O_2 dissolved in the water

   $${}^1/{}_2O_2 + 2H^+ + 2\,e^- \rightarrow H_2O$$

   The acid needed for this comes from CO_2 dissolved in the water, forming an acidic solution. Concurrently, dissolved O_2 combines with the Fe^{2+} ions as

   $${}^1/{}_2O_2 + 2Fe^{2+} + 5H_2O \rightarrow Fe_2O_3{\cdot}3H_2O + 4H^+$$

   This last compound is rust, a hydrated form of iron(III) oxide. Essentially the oxygen atom oxidized Fe^{2+} ions to Fe^{3+} ions.

   An interesting feature of the process is that these three reactions (two of them redox half-reactions) can each occur at a different spot on the metal surface. Thus corrosion (the first reaction) can be seen in one spot while rust formation (the third reaction) can be seen at another. It is believed that salt water catalyzes the process by helping to transport Fe^{2+} ions and electrons to different points along the metal surface.

5. Formation of stalagmites and stalactites in *Solubility and Precipitation* module.
6. Space scientists suggest the removal of carbon dioxide from the cabin atmosphere of a space mission by absorption in sodium hydroxide. Since the average human body breathes out about 925 g of carbon dioxide per day, for a six-passenger, 10-day expedition into space, 101 kg of sodium hydroxide would be required.

   $$CO_2(g) + 2NaOH(s) \rightarrow Na_2CO_3(aq) + H_2O(l)$$

7. If astronauts drank an average volume of 2 L of water a day, they would eliminate 2.4 L of water, with the additional amount coming from food metabolism. Scientists suggest that astronauts use lithium oxide (because of its low molar mass) to absorb the water. For the expedition hypothesized above, 240 kg of Li_2O would be needed.

   $$Li_2O(s) + H_2O \rightarrow 2LiOH(aq)$$

   You might try carrying through the stoichiometric calculations.
8. Fermentation, a not so simple chemical reaction, is the basis of the brewing and baking industries.

References

Module developed by Diana Doepken Rischling, Michael Pavelich, and Patricia Smith, the Colorado (South) team.

de Vos, W., and Verdonk, A. H. *Journal of Chemical Education.*

This series of articles entitled "A New Road to Reactions" is very valuable in teaching chemical reactions.

(March 1985). Part I: Chemical reactions to introduce and reinforce chemical concepts through a "guided discovery approach." *62*(3), 238-240.

(August 1985). Part II: Precipitation. *62*(8), 648-649.

(November 1986). Part III: Teaching the heat effect of reactions. *63*(11), 972-974.

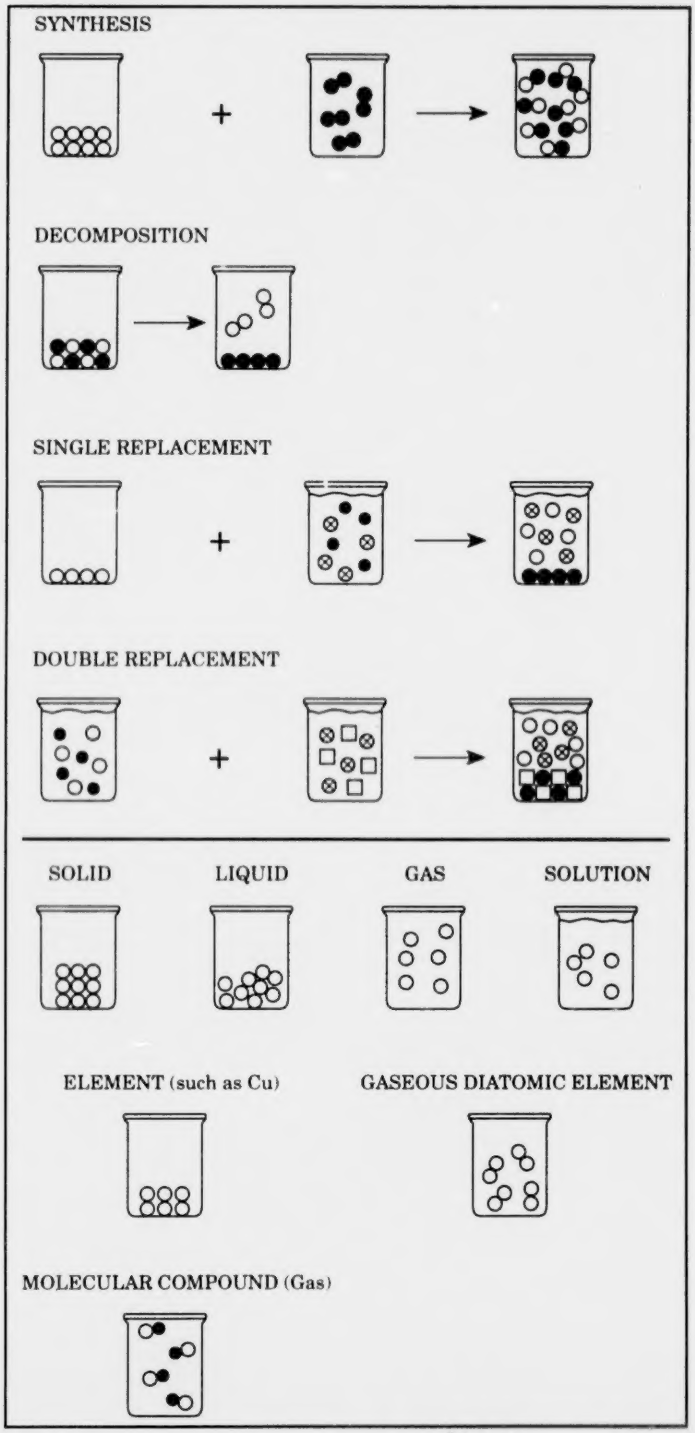
(August 1987). Part IV: The substance and its molecules. *64*(8), 692-694.

(December 1987). Part V: The elements and its atoms. *64*(12), 1010-1013.

Appendix

- **Transparency Masters**
  1. Solids, Liquids, Gases: Pictures in the Mind
  2. Word Search
- **Humor**

Solids, Liquids, Gases: Pictures in the Mind


SYNTHESIS
+
DECOMPOSITION
SINGLE REPLACEMENT
+
DOUBLE REPLACEMENT
+
SOLID
LIQUID
GAS
SOLUTION
ELEMENT (such as Cu)
GASEOUS DIATOMIC ELEMENT
MOLECULAR COMPOUND (Gas)

Word Search

S	Z	E	T	I	N	A	E	J	P	F	O	Q	I	W	W	P
F	Y	C	R	U	L	C	O	M	P	O	U	N	D	T	T	R
Z	G	N	E	X	Y	M	M	U	G	V	K	Q	E	C	E	E
U	Z	A	T	E	C	C	X	Y	A	B	P	W	D	Y	L	C
X	S	T	E	H	Y	X	N	O	I	T	A	U	Q	E	W	I
R	M	S	Q	X	E	Y	G	X	V	T	H	I	M	P	B	P
E	P	B	W	A	J	S	W	V	D	Y	W	E	V	A	X	I
A	L	U	K	H	E	D	I	C	Y	A	N	G	K	Q	U	T
C	Y	S	B	K	G	M	I	S	X	T	B	K	G	U	T	A
T	J	M	Q	X	J	C	X	B	S	A	M	N	N	E	J	T
I	F	G	M	I	X	T	U	R	E	K	U	C	U	O	N	E
O	R	M	Q	A	H	H	T	D	A	C	R	D	W	U	G	T
N	O	I	T	I	S	O	P	M	O	C	E	D	B	S	V	Q

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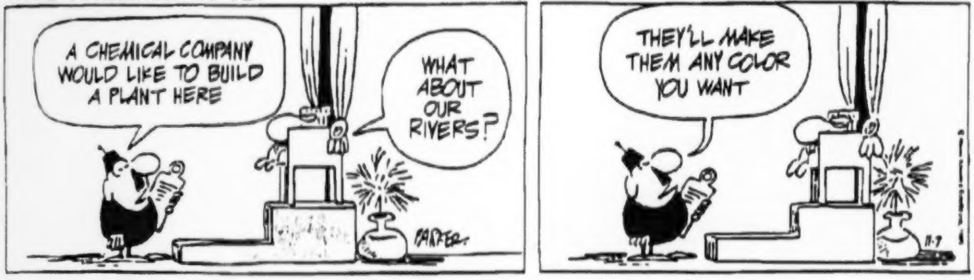
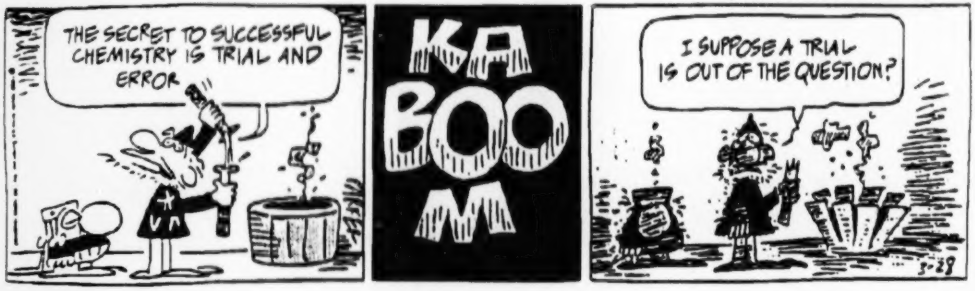
9. Symbolism used to represent reactions that occur.
10. Adjective for a water solution.

THE WIZARD OF ID


A CHEMICAL COMPANY WOULD LIKE TO BUILD A PLANT HERE
WHAT ABOUT OUR RIVERS?
THEY'LL MAKE THEM ANY COLOR YOU WANT
PARKER
11-7

By Brant Parker & Johnny Hart


THE SECRET TO SUCCESSFUL CHEMISTRY IS TRIAL AND ERROR
KA BOOM
I SUPPOSE A TRIAL IS OUT OF THE QUESTION?
3-28

By permission of Johnny Hart and NAS, Inc.

There's old bootsie using a Bunsen Burner same as I did.

CHEMTECH, August 1972, p. 510

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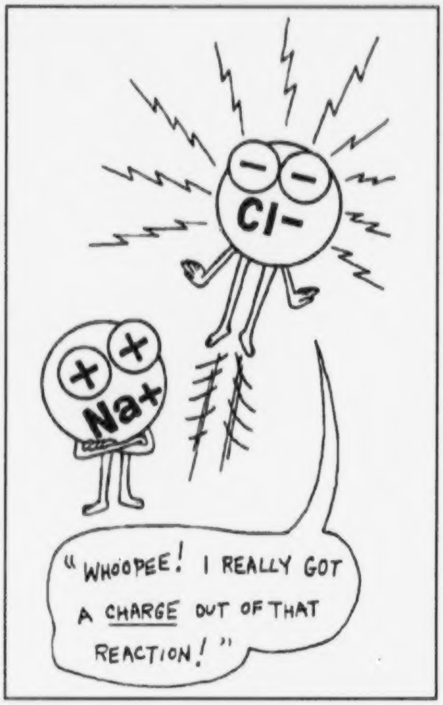
Cl-
Na+
"WHOOPEE! I REALLY GOT A CHARGE OUT OF THAT REACTION!"

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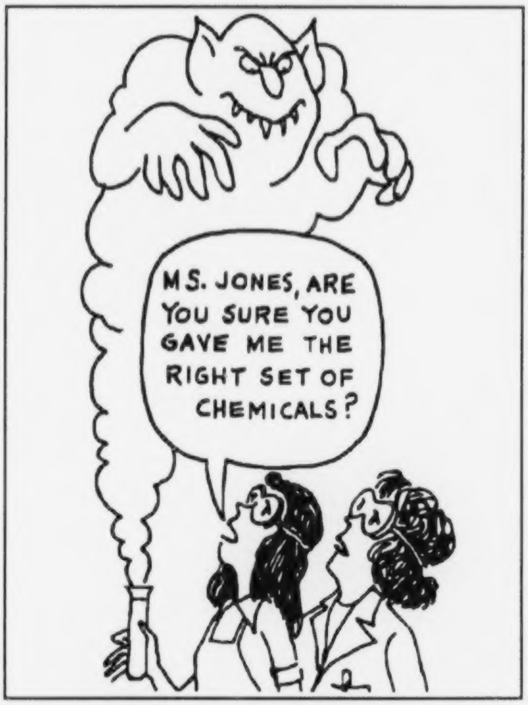
MS. JONES, ARE YOU SURE YOU GAVE ME THE RIGHT SET OF CHEMICALS?

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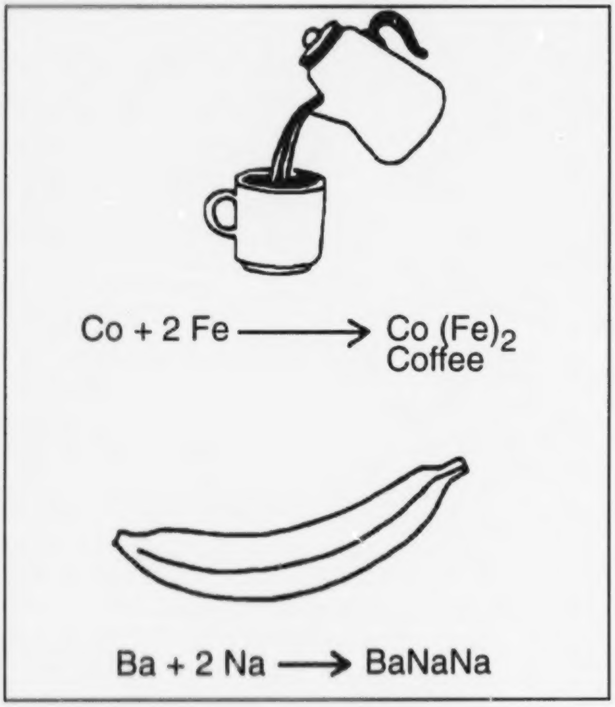
SOME SULFUR COMPOUNDS HAVE MILDLY UNPLEASANT ODORS!
uwe

CHEM 13 NEWS, February 1982, p. 12CHEM 13 NEWS, February 1982, p. 12. Reprinted with permission.


Co + 2 Fe ⟶ Co (Fe)₂
Coffee
Ba + 2 Na ⟶ BaNaNa

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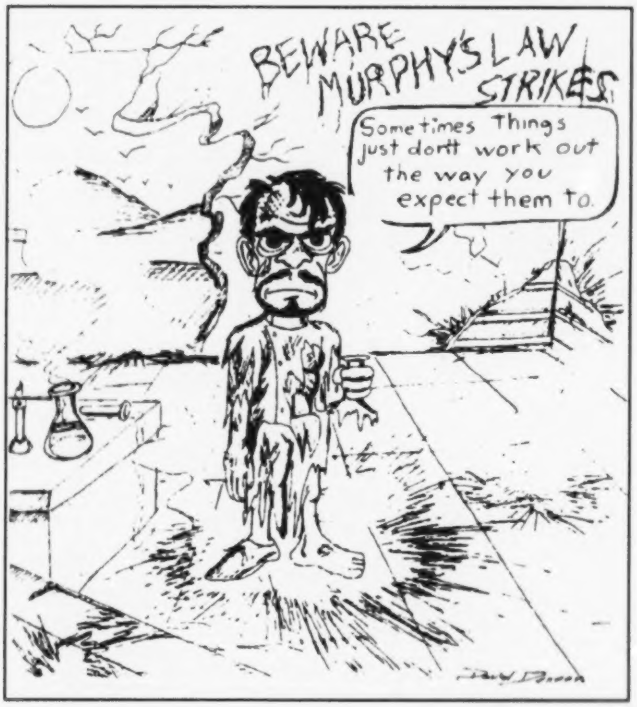
BEWARE MURPHY'S LAW STRIKES
Sometimes things just don't work out the way you expect them to.

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